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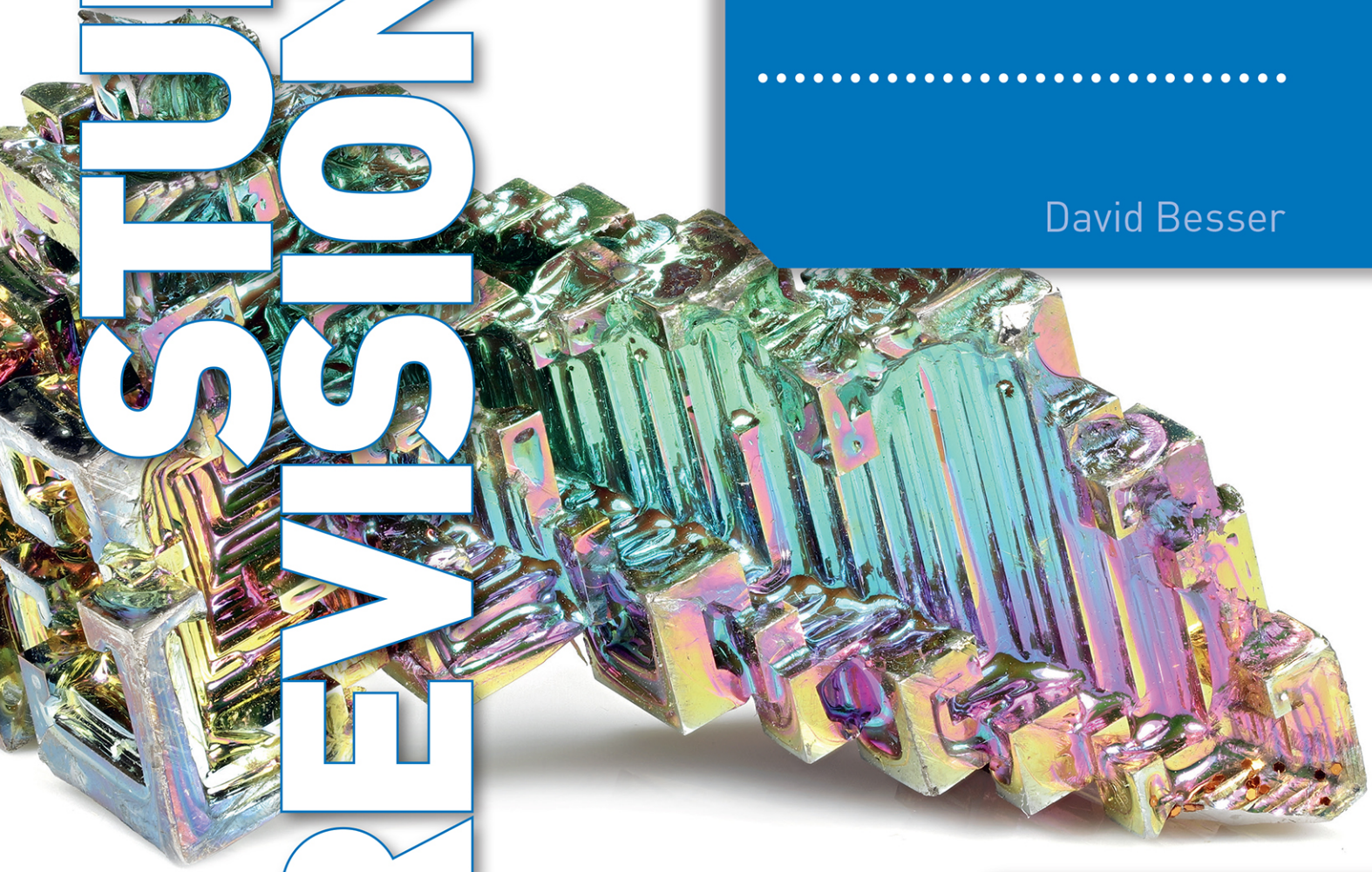
Chemistry

Third Edition

David Besser



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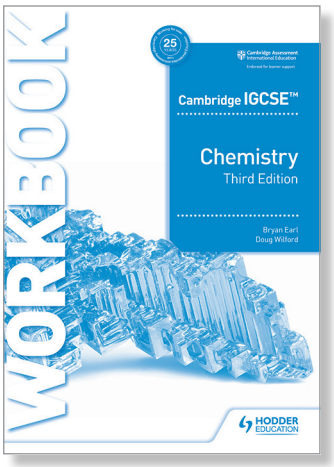

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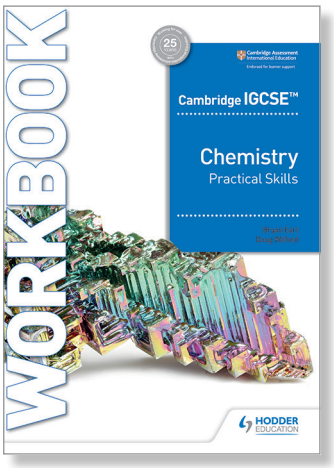

STUDY AND REVISION GUIDE

Working for over 25 YEARS WITH Cambridge Assessment International Education

Cambridge IGCSE™

Chemistry

Third Edition

David Besser

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Author dedication

Welcome to the world: Lucy Isabelle Besser and Finley Temple Williams

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Answers to exam-style questions:
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Introduction

Welcome to the Cambridge IGCSE™ Chemistry Study and Revision Guide. This book has been written to help you revise everything you need to know and understand for your Chemistry exam. Following the Chemistry syllabus, it covers all the key core and extended content and provides sample questions, as well as practice questions, to help you learn how to answer questions and to check your understanding.

How to use this book

Key objectives

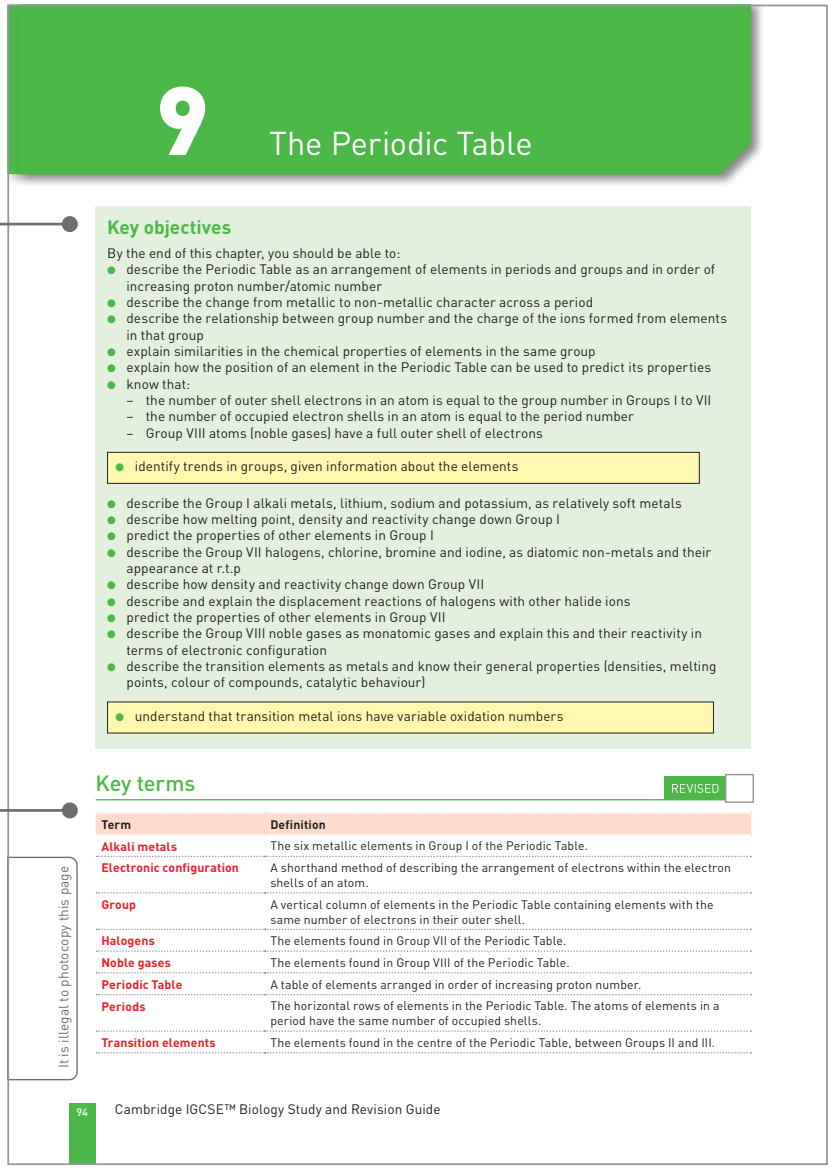
The key skills and knowledge covered in the chapter. You can also use this as a checklist to track your progress.

Key terms

Definitions of key terms you need to know from the syllabus.

Answers

Worked answers to the Exam-style questions can be found at **www.hoddereducation.co.uk/cambridgeextras**.



9 The Periodic Table

Key objectives

By the end of this chapter, you should be able to:
- describe the Periodic Table as an arrangement of elements in periods and groups and in order of increasing proton number/atomic number
- describe the change from metallic to non-metallic character across a period
- describe the relationship between group number and the charge of the ions formed from elements in that group
- explain similarities in the chemical properties of elements in the same group
- explain how the position of an element in the Periodic Table can be used to predict its properties
- know that:
  - the number of outer shell electrons in an atom is equal to the group number in Groups I to VII
  - the number of occupied electron shells in an atom is equal to the period number
  - Group VIII atoms (noble gases) have a full outer shell of electrons

- identify trends in groups, given information about the elements

- describe the Group I alkali metals, lithium, sodium and potassium, as relatively soft metals
- describe how melting point, density and reactivity change down Group I
- predict the properties of other elements in Group I
- describe the Group VII halogens, chlorine, bromine and iodine, as diatomic non-metals and their appearance at r.t.p
- describe how density and reactivity change down Group VII
- describe and explain the displacement reactions of halogens with other halide ions
- predict the properties of other elements in Group VII
- describe the Group VIII noble gases as monatomic gases and explain this and their reactivity in terms of electronic configuration
- describe the transition elements as metals and know their general properties (densities, melting points, colour of compounds, catalytic behaviour)

- understand that transition metal ions have variable oxidation numbers

Key terms REVISED

Term	Definition
Alkali metals	The six metallic elements in Group I of the Periodic Table.
Electronic configuration	A shorthand method of describing the arrangement of electrons within the electron shells of an atom.
Group	A vertical column of elements in the Periodic Table containing elements with the same number of electrons in their outer shell.
Halogens	The elements found in Group VII of the Periodic Table.
Noble gases	The elements found in Group VIII of the Periodic Table.
Periodic Table	A table of elements arranged in order of increasing proton number.
Periods	The horizontal rows of elements in the Periodic Table. The atoms of elements in a period have the same number of occupied shells.
Transition elements	The elements found in the centre of the Periodic Table, between Groups II and III.

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Teacher's comments

Feedback from a teacher showing what was good, and what could be improved.

Extended syllabus

Content for the Extended syllabus (supplemental material) is shaded yellow.

Sample questions

Exam-style questions for you to think about.

A hydrogen atom contains one electron. It can gain one electron or lose one electron to achieve a full outer shell.

a Write the symbol for the particle that forms when a hydrogen atom:
  i gains an electron [1]
  ii loses an electron. [1]

b Give two pieces of evidence that suggest hydrogen should be present in Group I of the Periodic Table rather than in Group VII. [2]

c Give one piece of evidence that suggests hydrogen should be present in Group VII of the Periodic Table rather than in Group I. [1]

Revision activity

There are many similar facts to learn in this chapter. If music helps you to concentrate, try playing the same song or tune every time you revise Group I, a different song for Group VII, and so on. (If you find background noise distracting, this isn't the method for you.)

Student's answer

$C_6H_{16} \rightarrow C_3H_8 + C_3H_6 + H_2$

Teacher's comment

The ending *–ane* indicates that hexane is an alkane and the general formula C_nH_{2n+2} should be used to deduce its formula. *Hex-* indicates that $n = 6$. The student begins with the incorrect formula for hexane, which makes it impossible to achieve the correct answer. H_2 is added as an attempt to 'balance' the equation.

Correct answer

Both of the following are acceptable answers because both produce an alkane and an alkene with the same number of carbon atoms:

$$C_6H_{14} \rightarrow C_3H_8 + C_3H_6$$

$$C_6H_{14} \rightarrow C_2H_6 + 2C_2H_4$$

Neither answer is more correct than the other.

If the question had specified a 1:1 mole ratio of the products, only the following would be correct:

$$C_6H_{14} \rightarrow C_3H_8 + C_3H_6$$

Skills

Where does an element belong?

Worked example

Sulfur has a proton number of 16.

State in which group and period of the Periodic Table sulfur is found.

Explain how you deduced your answers.

Answer

All atoms contain equal numbers of protons and electrons.

Therefore, a sulfur atom contains 16 electrons.

16 electrons give an electronic configuration of 2,8,6.

The group number is the same as the number of electrons in the outer shell.

Therefore, sulfur is in Group VI (6).

The period number is the number of shells that contain electrons.

Therefore, sulfur is in Period 3.

Exam-style questions

1 Use the following list of organic compounds to answer the questions that follow.
ethane ethene methane nylon poly(ethene)
Each substance can be used once, more than once or not at all.
For some question, you need to name only one substance. For others, there is more than one answer required.
Give the name of the substance or substances that:
  a are unsaturated [1]
  b are alkanes [2]
  c are formed by addition polymerisation [1]
  d contain a carbon–carbon double bond [1]
  e can act as a monomer [1]
  f are members of the same homologous series [2]
  g can be formed by hydrogenation of an alkene [1]
[Total: 9]

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Revision activities

Examples of strategies to help you revise effectively.

Student's answers

Typical student answers to see how the question might have been answered.

Skills

Key practical skills coverage in this guide will help you to consolidate your understanding of the practical work you have undertaken in your lessons and help you to describe and evaluate these skills effectively.

Key mathematical skills are covered to help you demonstrate these skills correctly.

Correct answer

Model student answers, corrected by the teacher's comment on the typical student answer.

Exam-style questions

Practice questions, set out as you would see them in the exam paper, for you to answer so that you can see what you have learned.

Exam breakdown

You will take three examinations at the end of your studies. If you have studied the Core syllabus content, you will take Paper 1 and Paper 3, and either Paper 5 or Paper 6. If you have studied the Extended syllabus content (Core and Supplement), you will take Paper 2 and Paper 4, and either Paper 5 or Paper 6.

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Paper 1: Multiple Choice (Core)	Paper 3: Theory (Core)
45 minutes	1 hour 15 minutes
40 marks	80 marks
40 four-option multiple-choice questions	Short-answer and structured questions
30% of your grade	50% of your grade

Paper 2: Multiple Choice (Extended)	Paper 4: Theory (Extended)
45 minutes	1 hour 15 minutes
40 marks	80 marks
40 four-option multiple-choice questions	Short-answer and structured questions
30% of your grade	50% of your grade

Paper 5: Practical Test	Paper 6: Alternative to Practical
1 hour 15 minutes	1 hour
40 marks	40 marks
Questions will be based on the experimental skills in Section 4	Questions will be based on the experimental skills in Section 4
20% of your grade	20% of your grade

Examination terms explained

The examination syllabus gives a full list of the terms used by examiners and how you are expected to respond.

Command word	Explanation
Calculate	Work out from given facts, figures or information. Give a numerical answer, generally showing the working out involved
Compare	Identify/comment on similarities and/or differences
Define	Give precise meaning
Describe	State the points of a topic / give characteristics and main features. An explanation is not required
Determine	Establish an answer using the information available
Evaluate	Judge or calculate the quality, importance, amount, or value of something
Explain	Set out purposes or reasons / make the relationships between things evident / provide why and/or how and support with relevant evidence
Give	Produce an answer from a given source or recall/memory
Identify	Name/select/recognise
Outline	Set out main points briefly, without going into detail
Predict	Suggest what may happen based on available information. You are not supposed to know the answer from memory, but to deduce it, usually from information in the question
Sketch	Make a simple freehand drawing showing the key features, taking care over proportions
State	Express in clear terms. No explanation is needed
Suggest	Apply knowledge and understanding to situations where there are a range of valid responses in order to make proposals / put forward considerations

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1 States of matter

Key objectives

By the end of this section, you should be able to:

- state the distinguishing properties of solids, liquids and gases
- describe the structure of solids, liquids and gases in terms of particle separation, arrangements and motion
- describe changes of state in terms of melting, boiling, evaporating, freezing and condensing
- describe the effects of temperature and pressure on the volume of a gas
- describe and explain diffusion in terms of kinetic particle theory
- explain changes in state in terms of kinetic particle theory, including the interpretation of heating and cooling curves
- explain in terms of kinetic particle theory the effects of temperature and pressure on the volume of a gas
- describe and explain the effect of relative molecular mass on the rate of diffusion of gases

Key terms

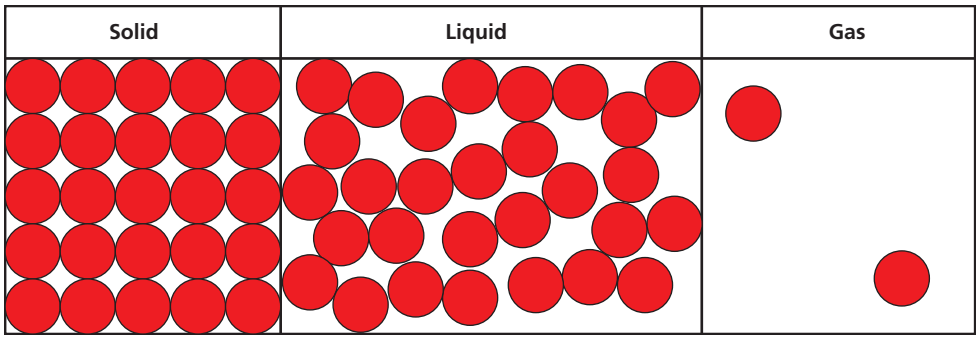
REVISED

Term	Definition
Boiling point	The temperature at which the pressure of the gas created above a liquid equals atmospheric pressure.
Condensation	The change of a gas into a liquid. This process is accompanied by the evolution of heat.
Diffusion	The process by which different substances mix as a result of the random motions of their particles.
Evaporation	A process occurring at the surface of a liquid involving the change of state of a liquid into a gas at a temperature below the boiling point. When a solution is heated, the solvent evaporates and leaves the solute behind.
Freezing point	The temperature at which a substance freezes. This has the same value as the melting point.
Melting point	The temperature at which a solid begins to turn into a liquid. Pure substances have a sharp melting point.

1.1 Solids, liquids and gases

REVISED

The differences between solids, liquids and gases in terms of particle arrangement and particle separation are shown in Figure 1.1.



▲ **Figure 1.1 Particle arrangements in a solid, a liquid and a gas. Note how most of the particles in the liquid are touching.**

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The differences in the properties of solids, liquids and gases, along with the explanations for these differences, based on kinetic theory, are shown in Table 1.1.

▼ **Table 1.1 Properties of solids, liquids and gases**

		Surface boundary	Shape
Solids	Property	Solids have a surface boundary.	Solids have a fixed shape.
	Explanation	Strong forces of attraction between particles in a solid prevent particles from escaping.	Strong forces of attraction between particles in solids mean that the particles are held together in a fixed shape. The particles vibrate about fixed positions but do not move from place to place.
Liquids	Property	Liquids have a surface boundary.	Liquids take the shape of the container that they are present in.
	Explanation	The forces of attraction between the particles in a liquid are strong enough to prevent the majority of the liquid particles from escaping and becoming a gas.	The forces of attraction between particles in a liquid are weaker than in solids. The particles slowly move from place to place, meaning that a liquid can change its shape to fit the container.
Gases	Property	Gases have no surface boundary.	Gases fill the container they are held in. They have no fixed shape.
	Explanation	Gas particles move at high speeds. The particles have only very small forces of attraction between them.	The forces of attraction between gas particles are extremely weak. The gas particles move at very high speeds therefore gases move to fill the container.

Revision activity

Make a table of your own to show the key information from this section. Decide which headings you need and use notes instead of complete sentences. You might want to include diagrams in some cells. It is important the table is personal to you.

1.2 Kinetic theory

REVISED

When heat energy is given to a solid, the heat energy causes the particles to vibrate faster and faster about a fixed position. Eventually, the particles have enough energy for **melting** to occur. At the **melting point**, the particles have enough energy to overcome the forces of attraction between them when they are in the solid. The ordered arrangement breaks down as the solid turns into a liquid.

There is no further increase in temperature until all the solid has turned into a liquid – that is, the ordered arrangement has completely broken down.

After this, the energy given to the particles causes them to move faster from place to place until they have enough energy for **boiling** to occur. At the **boiling point**, the particles have enough energy to almost completely overcome the attraction between them when they are in the liquid. They then move as far away from each other as possible.

Again, there is no increase in temperature until all the liquid has turned into a gas.

If a gas is heated, the particles gain more and more energy and move at increasing speeds – the temperature of the gas increases.

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A liquid can become a gas by **evaporation**. This is not the same as boiling.

- Evaporation only occurs on the surface of a liquid, whereas boiling occurs throughout the liquid.
- Boiling only takes place at the **boiling point** of a liquid, but evaporation occurs at temperatures below the boiling point (as well as at the boiling point).

Puddles of water evaporate on a sunny day even though the water in the puddles does not reach 100°C. The water on the surface of the puddle turns into water vapour at temperatures well below the boiling point of water.

1.3 Changes of state

REVISED

Figure 1.2 summarises the changes in state between solids, liquids and gases.

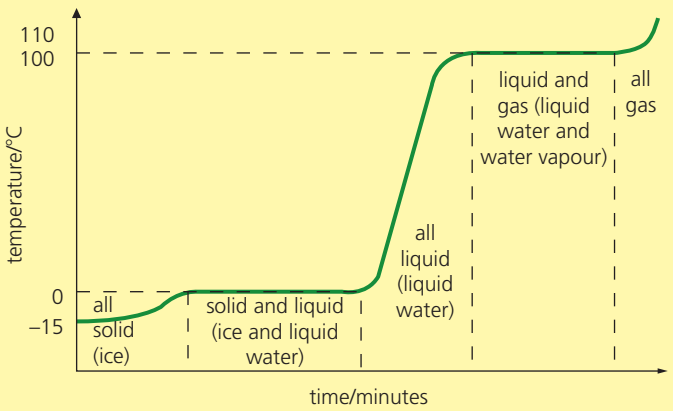
▲ **Figure 1.2 Changes of state**

Heating and cooling curves

Heating curve

A heating curve shows how a solid changes state when the temperature is gradually increased.

- The process shown in Figure 1.3 begins with ice at a temperature below 0°C. The temperature gradually increases until it reaches 0°C, which is the **melting point** of ice.
- At this point, ice and water exist together. The temperature does not change until all the ice has changed into water, which is why the graph line is horizontal. A sharp melting point (at one specific temperature) is an indication that a solid is pure.
- At the melting point, the particles have gained enough energy to overcome the forces of attraction that keep them in position in the solid. When the line is horizontal, the temperature is constant – the particles gain no more energy until all the attractions between them that exist in the solid are overcome.



▲ **Figure 1.3 Heating curve showing the change of temperature against time for the change from ice at –15°C to water to water vapour (steam)**

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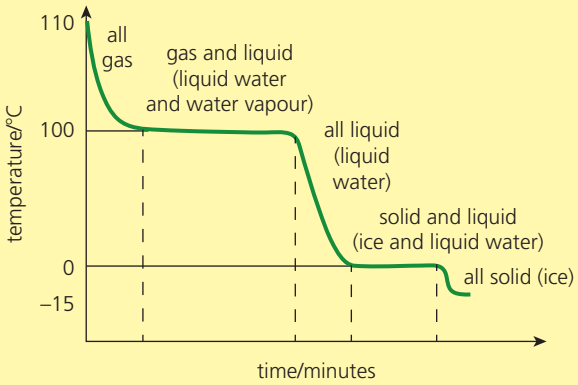
- The temperature then begins to increase again until it reaches 100°C, which is the **boiling point** of water.
- The temperature does not increase again until all the water has changed into water vapour. This is why the line is horizontal for a second time.
- At the boiling point, the particles have gained enough energy to overcome the forces of attraction between them. The particles stop gaining energy until there is almost no attraction between them.
- When all the water has boiled, the temperature begins to rise again as the particles in the gaseous state gain more energy.

Cooling curve

A similar curve results when a gas is cooled gradually until it forms a solid. This is known as a cooling curve.

Figure 1.4 shows how water vapour changes state when the temperature is gradually decreased.



▲ **Figure 1.4 Cooling curve**

- At the start, the water vapour is at a temperature above 100°C. The temperature gradually decreases until it reaches 100°C, which is the **boiling point** of water.
- At this point water vapour and water exist together. The temperature does not change until all the water vapour has changed into liquid water, which is why the graph line is horizontal.
- At the boiling point, the particles stop losing energy as forces of attraction between them form. The temperature does not decrease again until the forces of attraction between water particles in the liquid state have been fully formed.
- The temperature then begins to decrease again until it reaches 0°C, which is the **freezing point** of water.
- The temperature does not change again until all the liquid water has changed into ice, which is why the line is horizontal for a second time.
- At the freezing point, the particles stop losing energy as forces of attraction needed to hold them in position in a solid form between them. The temperature does not further decrease until these forces have been fully formed.
- When all the liquid water has frozen, the temperature begins to fall again as the particles in the solid state lose more energy.

Revision activity

The tick boxes in this book help you to record the topics you have revised. Combine this with a revision diary, in which you make a note of the sections you need to review again or concepts you need to ask your teacher about. Remember to write down what went well in each session, too.

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1.4 Effects of temperature and pressure on the volume of a gas

REVISED

As the temperature of a gas increases, the volume of the gas increases proportionally if the pressure remains constant.

- When the temperature of a gas increases, the particles gain energy.
- The particles therefore collide with the walls of the container more often and with greater force.
- If the container does not have a fixed volume (if it is, for example, a gas syringe or a balloon), the volume increases in order to maintain a constant pressure.

As the pressure of a gas increases, the volume of the gas decreases proportionally if the temperature remains constant.

- When the pressure of a gas increases, the particles move closer together.
- This means that the same number of particles occupy a smaller space, which is why the volume of gas decreases.

1.5 Diffusion

REVISED

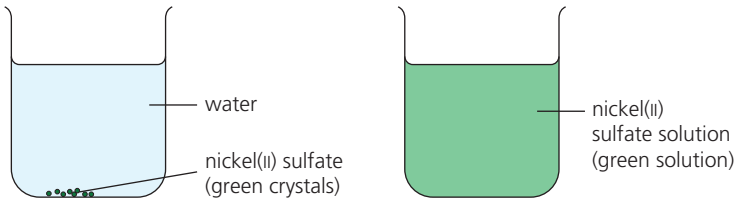
Particles in solids do not move from one place to another – they only vibrate. Particles in liquids move slowly and particles in gases move much more quickly.

Diffusion can be demonstrated experimentally in liquids and in gases.

Diffusion in liquids

If crystals of a coloured solid, such as nickel(II) sulfate, are placed in a liquid, such as water, the colour of the nickel(II) sulfate spreads throughout the liquid in a matter of days, producing a solution with a uniform colour.

This is because the particles (ions) in nickel(II) sulfate move randomly from where there are a lot of them (high concentration) to where there are fewer of them (low concentration).



▲ **Figure 1.5 Diffusion in a liquid**

Diffusion in gases

Diffusion happens much faster in gases than in liquids. This is because gas particles move more quickly than liquid particles.

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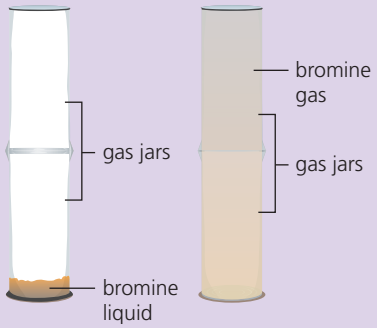
Skills

Diffusion in bromine

If bromine liquid is placed in the bottom of a gas jar with another gas jar on top, as shown in Figure 1.6, the liquid evaporates and the brown colour of the bromine gas fills both jars after a short time.



▲ **Figure 1.6 Diffusion in a gas**

This experiment shows that the particles (molecules) in bromine move randomly from where there are a lot of them (high concentration) to where there are fewer of them (low concentration).

Breathing in bromine gas causes respiratory problems and dizziness. Getting bromine liquid or gas on your skin can cause skin irritation and burns. For these reasons, bromine is only used in demonstration experiments by a teacher. The teacher should wear protective gloves and carry out this experiment in a fume cupboard.

Different gases diffuse at different rates when at the same temperature because of differences in their relative molecular masses. In fact, the rate at which a gas diffuses is inversely related to the relative molecular mass of the gas.

Lighter particles travel faster and further than heavier molecules. Therefore, particles in gases with lower relative molecular mass will diffuse faster than particles in gases with higher relative molecular mass.

It is important to remember that the inverse relationship between rate of diffusion and relative molecular mass only applies to gases and *not* to solids or liquids.

Revision activity

Use your revision diary to help you work out which techniques help you to learn. When you record what went well in a session, make a note about *how* you revised the topic. Did you work with a friend? Did you put information into a table or draw a mind map? Then next time, try using that method to revise the sections you find difficult.

Sample questions

REVISED

1 A compound has a melting point of –30°C and a boiling point of 85°C. Give the physical state of the compound at 25°C. Explain your answer. [2]

Student's answer

Liquid because the melting point is below 25°C and the boiling point is above 25°C.

Teacher's comments

The student's answer is correct.

There are two things to look for in a question of this type.

- Pay attention to the negative sign. If a student ignores this and takes the melting point as 30°C, which is higher than 25°C, they will think the compound is a solid.
- Use all the information. A student who states only that the melting point is below 25°C without mentioning the boiling point will get some credit but not full marks.

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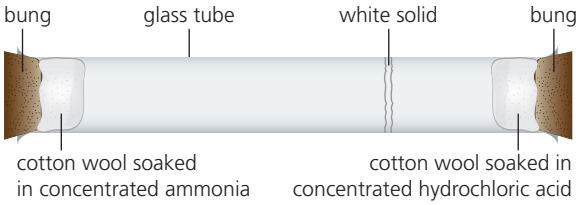
2 When the apparatus shown below is set up, concentrated ammonia releases ammonia gas, NH_3, and concentrated hydrochloric acid releases hydrogen chloride gas, HC*l*.



When ammonia gas reacts with hydrogen chloride gas, a white solid is produced according to the equation:

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

a Give the name of the white solid.
b Name the process by which the two gases move through the glass tube.

c Explain why the white solid forms nearer the concentrated hydrochloric acid end of the glass tube rather than the ammonia solution end.

Student's answers

a *ammonia chloride*
b *diffusion*
c *Ammonia and hydrogen chloride diffuse through the glass tube. Ammonia is lighter than hydrogen chloride, so it diffuses faster and the gases meet nearer the hydrochloric acid end.*

Teacher's comments

a Ammonium compounds are often mistakenly referred to as *ammonia* compounds. Similarly, ammonia is often referred to as *ammonium*. Make sure you are aware of the difference between ammonia gas, NH_3, and the ammonium ion, NH_4^+,which is part of all ammonium salts.
b This is the correct answer.

c This answer would gain very little (if any) credit. Answers must refer to ammonia and hydrogen chloride molecules and state that ammonia has a smaller relative molecular mass than hydrogen chloride. You should calculate relative molecular masses using relative atomic masses given in the Periodic Table if they are not provided in the question (see Section 4.1).

Correct answers

a ammonium chloride
b diffusion

c Molecules of ammonia and hydrogen chloride diffuse through the glass tube. Ammonia, NH_3, has a lower relative molecular mass (17) than hydrogen chloride, HC*l* (36.5), so ammonia molecules diffuse faster than hydrogen chloride molecules. Therefore, the gases meet and react nearer the cotton wool soaked in hydrochloric acid.

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Exam-style questions

1 A substance has a melting point of 85°C and a boiling point of 180°C. Give the physical state of the substance at 50°C. Explain your answer. [Total: 2]

2 Use the letters A, B, C and D to answer the questions under the table.

Substance	Distance between particles	Arrangement of particles	Movement of particles
A	Very far apart	Ordered	Vibrate about fixed position
B	Fairly close together	Irregular	Move slowly
C	Very close together	Ordered	Vibrate about fixed position
D	Very far apart	Random	Move at high speeds

Give the letter of the substance that is:

a a solid [1]
b a liquid [1]
c a gas [1]
d unlikely to represent a real substance [1]

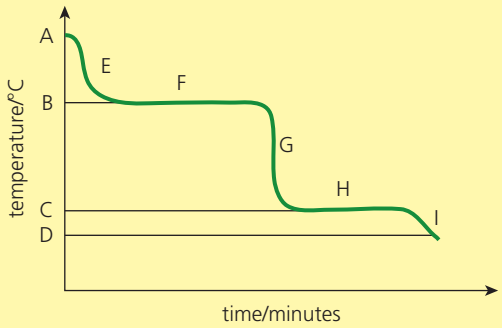
[Total: 4]

3 State the word that represents the following changes:

a the change of a gas into a liquid [1]
b the process by which different substances mix as a result of the random motions of their particles [1]
c the process that occurs when a liquid turns into a solid [1]
d the process that occurs on the surface of a liquid when the liquid turns into a gas at a temperature below the boiling point [1]

[Total: 4]

4 A cooling curve for solid **X** is shown below.



a Give the letter for the temperature that shows the freezing point of **X**. [1]
b Give the letter that represents the part of the curve where **X** exists as a gas only. [1]
c Describe the changes in particle separation, arrangement, motion and forces of attraction that occur when **X** changes from a liquid to a solid. [4]

[Total: 6]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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2 Atoms, elements and compounds

Key objectives

By the end of this section, you should be able to:

- describe the differences between elements, compounds and mixtures
- interpret and use symbols for given atoms
- define the molecular formula of a compound as the number and type of different atoms in one molecule
- deduce the formula of a simple compound from the relative numbers of atoms present in a model or a diagram of a molecule
- write word equations and symbol equations (including state symbols) to show how reactants form products
- define oxidation as oxygen gain and reduction as oxygen loss
- identify redox reactions as reactions involving gain and loss of oxygen

- define an oxidising agent and a reducing agent

- describe the structure of an atom as a central nucleus containing neutrons and protons surrounded by electrons in shells
- state the relative charges and relative masses of a proton, a neutron and an electron
- define proton number (atomic number) and mass number (nucleon number)
- determine the electronic configuration of atoms and ions with proton number 1 to 20
- describe the formation of positive ions, known as cations, and negative ions, known as anions
- define isotopes

- state that isotopes of the same element have the same chemical properties and give the reason for this
- calculate the relative atomic mass (A_r) of an element from the relative masses and abundances of its isotopes

Key terms

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Term	Definition
Anion	A negative ion.
Atom	The smallest part of an element that can exist as a stable entity. It has a central nucleus containing neutrons and protons surrounded by electrons in shells. An atom contains equal numbers of protons and electrons.
Cation	A positive ion.
Chemical change	A permanent change in which a new substance is formed.
Compound	A substance formed by the chemical combination of two or more elements in fixed proportions.
Diatomic molecule	A molecule containing two atoms.
Element	A substance that cannot be further divided into simpler substances by chemical methods.
Ion	An atom or group of atoms that has either lost one or more electrons, making it positively charged, or gained one or more electrons, making it negatively charged.
Isotopes	Different atoms of the same element that have the same number of protons but different numbers of neutrons.
Mass number (nucleon number)	The total number of protons and neutrons found in the nucleus of an atom, symbol A.
Mixture	Two or more substances mixed together that can be separated by physical means.
Molecule	A group of atoms covalently bonded together.
Monatomic molecule	A molecule which consists of only one atom.
Oxidation	Gain of oxygen.

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Term	Definition
Oxidising agent	A substance that oxidises another substance and is itself reduced.
Proton number (atomic number)	The number of protons in the nucleus of an atom, symbol *Z*.
Redox reaction	A reaction which involves simultaneous oxidation and reduction.
Reducing agent	A substance that reduces another substance and is itself oxidised.
Reduction	Loss of oxygen.
Relative atomic mass (A_r)	The average mass of the isotopes of an element compared to 1/12 of the mass of an atom of ^{12}C.

2.1 Elements

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The Periodic Table (see Chapter 4) consists of **elements** only. Each element is made up of only one type of **atom** and is represented by a chemical symbol.

Elements *cannot* be decomposed into anything simpler by chemical methods. Note that *smaller* is not the same as *simpler*. For example, a piece of sulfur can be broken with a hammer into several smaller pieces of sulfur, but this is not breaking it into anything simpler. The act of breaking with a hammer is a physical method and not a chemical method. Thus, sulfur is an element.

Elements are classified as metals and non-metals as shown in Table 2.1.

▼ **Table 2.1 Classification of elements**

Property	Metal	Non-metal
Physical state at room temperature	Solid (except mercury)	Solid, liquid (bromine only) or gas
Malleability	Good	Poor (usually soft or brittle)
Ductility	Good	Poor (usually soft or brittle)
Appearance	Shiny (lustrous)	Usually dull
Melting point/boiling point	Usually high	Usually low
Density	Usually high	Usually low
Conductivity (electrical and thermal)	Good	Poor (except graphite)

2.2 Compounds

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Compounds have a chemical formula which shows them to contain two or more elements which are chemically combined in fixed proportions.

Although compounds are made of more than one element, they are still pure substances because they have a constant composition – every **molecule** is the same.

Examples of compounds are:

- sodium chloride, $NaCl$
- carbon dioxide, CO_2
- copper(II) nitrate, $Cu(NO_3)_2$

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Skills

Formulae

The formula of a compound shows the proportions of each element. For example, iron(II) sulfide has the formula FeS. This means that all samples of iron(II) sulfide contain iron and sulfur in the proportion of 1 atom of iron to 1 atom of sulfur.

Worked example 1

Write down the ratio of atoms in lead(II) nitrate, $Pb(NO_3)_2$.

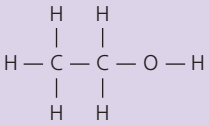
Answer

Multiplying out the brackets gives PbN_2O_6.

Therefore, the ratio of Pb:N:O is 1:2:6.

Worked example 2

Use Figure 2.1 to deduce the molecular formula of ethanol.



▲ **Figure 2.1 A molecule of ethanol**

Answer

The molecule contains 2 carbon atoms, 6 hydrogen atoms and 1 oxygen atom.

Therefore, the molecular formula is C_2H_6O.

Chemical changes and word equations

Chemical changes or **chemical reactions** are changes in which new chemical substances are produced.

Word equations give the names of the reactants which take part in a chemical reaction and the products that are made in the reaction.

In the following reaction:

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

we say the magnesium is **oxidised** because it *gains* oxygen.

Oxygen causes the **oxidation**. Therefore, oxygen is the **oxidising agent**.

Reduction is the opposite of oxidation – it is the *loss* of oxygen. In the following reaction:

copper oxide + hydrogen → copper + water

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$$

copper(II) oxide is reduced because it loses oxygen.

In this reaction, hydrogen gains oxygen – it is oxidised.

Hydrogen causes the reduction. Therefore, hydrogen is the **reducing agent**.

It follows that oxidation and reduction always occur at the same time.

A reaction in which oxidation and reduction both occur is known as a **redox reaction**.

Symbol equations give the correct formulae of the reactants and products in a reaction. Symbol equations are balanced when the number of atoms of each element is the same on both sides of the equation.

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Skills

Balancing symbol equations

1. Write down the word equation. (This can be omitted with experience.)
2. Write down the correct formulae of the reactants and products.
   Take care not to use incorrect formulae (e.g. H instead of H_2 or O instead of O_2) or to change formulae. In Worked example 1 below, changing H_2O into H_2O_2 would make the number of atoms of each element the same on both sides, but H_2O_2 is the formula for hydrogen peroxide not water.
3. Count the number of atoms of each element on both sides.
4. If the number of atoms of each element is not the same on both sides, put numbers *in front of* the formulae so that the number of atoms of each element is the same on both sides.
5. Put state symbols after the formulae:
   (s) = solid, (*l*) = liquid, (g) = gas, (aq) = aqueous solution. This can be done after Steps 2 or 3 if preferred.

Worked example 1

Word equation:

hydrogen + oxygen → water

Unbalanced equation:

$$H_2 + O_2 \rightarrow H_2O$$

Number of atoms of each element on both sides:

H 2	H 2
O 2	O 1

As the number of atoms of oxygen is not the same on both sides, the first step is to put **2** in front of H_2O. This multiplies everything that comes after it.

$$H_2 + O_2 \rightarrow \mathbf{2}H_2O$$

Number of atoms of each element on both sides:

H 2	H **4**
O 2	O **2**

In balancing the oxygen, we have unbalanced the hydrogen. Therefore, we need to put **2** in front of H_2. The equation is now balanced and state symbols can be inserted:

$$\mathbf{2}H_2(g) + O_2(g) \rightarrow \mathbf{2}H_2O(l)$$

Number of atoms of each element on both sides:

H **4**	H 4
O 2	O 2

Worked example 2

Word equation:

aluminium + chlorine → aluminium chloride

Unbalanced equation:

$$Al + Cl_2 \rightarrow AlCl_3$$

Number of atoms of each element on both sides:

Al 1	*Al* 1
Cl 2	*Cl* 3

The aluminium is balanced. To balance the chlorine, we put **2** in front of $AlCl_3$ and **3** in front of Cl_2:

$$Al + \mathbf{3}Cl_2 \rightarrow \mathbf{2}AlCl_3$$

Number of atoms of each element on both sides:

Al 1	*Al* **2**
Cl **6**	*Cl* **6**

The aluminium is now unbalanced, so we must put **2** in front of *Al* to balance the equation:

$$\mathbf{2}Al(s) + \mathbf{3}Cl_2(g) \rightarrow \mathbf{2}AlCl_3(s)$$

Number of atoms of each element on both sides:

Al **2**	*Al* 2
Cl 6	*Cl* 6

2.3 Mixtures

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Mixtures contain two or more elements and/or compounds in variable proportions. Mixtures do not have a chemical formula.

If a compound is present in an aqueous solution, the aqueous solution is a mixture because it contains two substances which are not chemically combined. For example, sodium hydroxide solution (also referred to as aqueous sodium hydroxide) contains sodium hydroxide and water.

A mixture containing two substances does not necessarily contain equal amounts of each substance. If we had a mixture of salt and sand which contained equal amounts of each substance and we added more salt to it, it would still be called a mixture of salt and sand. A mixture of salt and sand can contain more salt than sand, or more sand than salt, or equal

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amounts of salt and sand. This is different from the composition of a compound, as shown in the case of iron(II) sulfide above.

Air is another mixture (see Chapter 11). It does not have a chemical formula because it contains several chemical substances as opposed to one substance. Air contains nitrogen and oxygen with smaller amounts of other gases, such as water vapour, carbon dioxide and argon. Polluted air may also contain other gases, such as carbon monoxide, sulfur dioxide and nitrogen dioxide.

Air from different places has different percentages of its constituent gases. For example, the amounts of pollutant gases are lower in the countryside than in industrial areas. However, the different samples are all called air even though the composition of the mixture can vary.

2.4 Inside atoms

REVISED

Atoms are made from smaller particles called protons, neutrons and electrons. The protons and neutrons exist in the centre of the atom in a dense region called the nucleus. The electrons move around the nucleus and exist in electron shells at increasing distances from the nucleus.

Make sure you learn the information in Table 2.2. You need to know the differences between the **relative mass** and **relative charge** of a proton, neutron and electron.

▼ **Table 2.2 The properties of protons, neutrons and electrons**

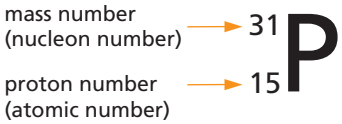
Particle	Relative mass/atomic mass units	Relative charge
Proton	1	+1
Neutron	1	0
Electron	1/1837	–1

Atoms are often represented as shown in Figure 2.2.

The **proton number** or **atomic number** is the number of protons in one atom of an element. As atoms do not have a charge, the number of protons in an atom is always equal to the number of electrons.

The **mass number** or **nucleon number** is the number of neutrons and protons added together in one atom of an element.

It is a good idea to remember that the mass number is always larger than the proton number (with the exception of the most abundant **isotope** of hydrogen, for which both numbers are 1).



▲ **Figure 2.2 Phosphorus as shown in a Periodic Table. In some textbooks, the two numbers may be reversed.**

Skills

Calculating the number and type of particles in an atom

proton number = number of protons in one atom
= number of electrons in one atom

number of neutrons = nucleon number – proton number

Worked example

Calculate the number of protons, electrons and neutrons in one phosphorus atom using the information in Figure 2.2.

Answer

number of protons = proton number = 15

number of electrons = number of protons = 15

number of neutrons = nucleon number – proton number
= 31 – 15 = 16

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Ions

Ions are atoms (or groups of atoms) that have gained or lost an electron or electrons.

Positive ions (**cations**) are formed when atoms or groups of atoms lose an electron or electrons. They are positively charged because the number of protons is larger than the number of electrons. The number of positive charges is equal to the number of electrons that are lost when they form.

Negative ions (**anions**) are formed when atoms or groups of atoms gain an electron or electrons. They are negatively charged because the number of electrons is larger than the number of protons. The number of negative charges is equal to the number of electrons that are gained when they form.

Skills

Calculating the number and type of particles in an ion

Worked example 1

Calculate the number of protons, neutrons and electrons in a Cu^{2+} ion.

Cu has a nucleon number of 63.

Answer

- If the proton number is not given in the question, you should use a Periodic Table to find it.
  Copper has a proton number of 29. This means that all copper atoms and ions contain 29 protons.
- The number of neutrons = nucleon number – proton number
  = 63 – 29 = 34.
- All positive ions have more protons than electrons.
  The charge on a positive ion = number of protons – number of electrons.

  For Cu^{2+} the charge = 29 – number of electrons
  2 = 29 – number of electrons
  Number of electrons = 29 – 2 = 27

Worked example 2

A particle contains 13 protons, 14 neutrons and 10 electrons.

a Give the symbol of the element that forms this particle.
b Deduce the mass number of the particle.
c Deduce the charge on the particle.

Answers

a The symbol of the element depends only on the number of protons in the particle. The element can be identified from the Periodic Table. Aluminium atoms and ions have 13 protons, so the symbol is *Al*.
b Mass number = number of protons + number of neutrons
= 13 + 14 = 27
c This ion contains more protons than electrons. Therefore, it is a positive ion.
Size of positive charge = number of protons – number of electrons
= 13 – 10 = 3+

Isotopes

Isotopes are atoms of the same element containing the same number of protons but different numbers of neutrons.

Examples of isotopes of argon are shown in Table 2.3.

▼ **Table 2.3 Isotopes of argon**

Isotope	Number of protons	Number of neutrons	Number of electrons
$^{40}_{18}Ar$	18	(40 – 18) = 22	18
$^{38}_{18}Ar$	18	(38 – 18) = 20	18
$^{36}_{18}Ar$	18	(36 – 18) = 18	18

Isotopes of the same element all have the same number of electrons.

Isotopes of the same element all have the same chemical properties because they have the same electronic configuration (see Chapter 9).

Revision activity

It is easy to confuse the terms *isotope* and (structural) *isomer* (see Chapter 12). Write key words like these on one side of a card – one word per card – and the definitions on the other side. Learn the words and definitions. Then use the cards to test yourself or a friend.

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Relative atomic mass (A_r)

Relative atomic mass (A_r) is the average mass of the isotopes of an element compared to 1/12 of the mass of an atom of ^{12}C.

Skills

Calculating A_r

The relative atomic mass of an element can be determined from the relative masses (mass numbers) and abundances of its isotopes.

Worked example

Use the information in Table 2.4 to calculate the relative atomic mass (A_r) of chlorine.

▼ **Table 2.4 Abundance of chlorine isotopes**

Relative mass (mass number)	Percentage abundance
35	75
37	25

Answer

The percentage abundances mean that if we consider 100 chlorine atoms:

- 75 have mass of 35 relative mass units each
- 25 have mass of 37 relative mass units each

Therefore, the total mass of 100 chlorine atoms is:

$$(35 \times 75) + (37 \times 25) = 3550$$

average mass of 1 chlorine atom = total mass of 100 atoms ÷ 100

$$= 3550 \div 100 = 35.5$$

Therefore, the relative atomic mass (A_r) of chlorine = 35.5.

The arrangement of electrons in an atom

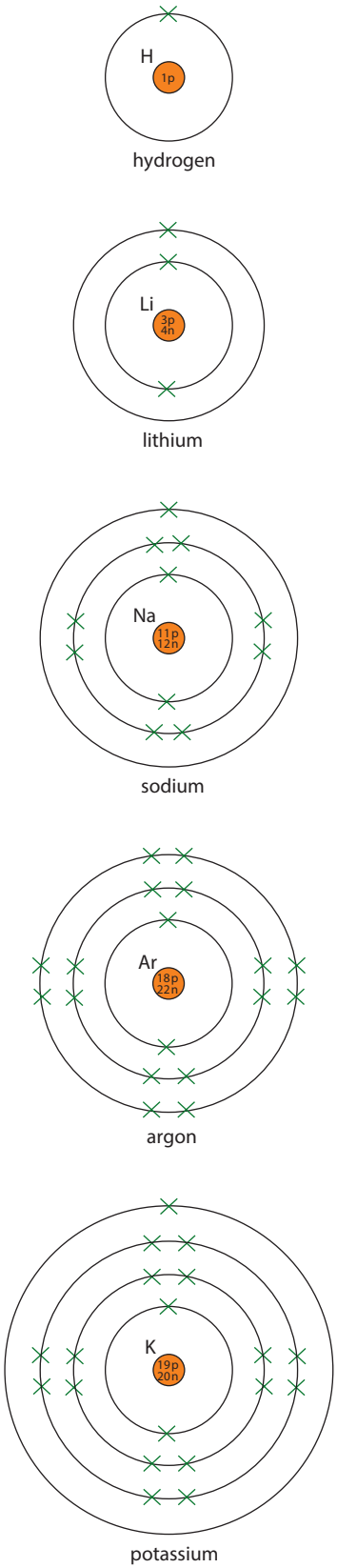
Electrons are arranged in electron shells at increasing distances from the nucleus. These shells can hold up to a maximum number of electrons, as shown in Table 2.5.

▼ **Table 2.5 Maximum number of electrons per shell**

Shell number	Maximum number of electrons
1	2
2	8
3	8*

*Shell 3 can, in fact, hold up to 18 electrons, but this does not need to be considered at this level.

These arrangements can be shown as a list of numbers, as in Table 2.6, or as diagrams, as in Figure 2.3.



▲ **Figure 2.3 Electronic configurations of hydrogen, lithium, sodium, argon and potassium**

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▼ **Table 2.6 Electronic configurations of helium, carbon, oxygen and phosphorus**

Element	Number of electrons	Electronic configuration
Helium, He	2	2
Carbon, C	6	2,4
Oxygen, O	8	2,6
Phosphorus, P	15	2,8,5

The structure of atoms affects the chemical properties of elements. You can find out more about this in Chapter 9.

Revision activity

Use a Periodic Table to help you draw a table showing the number, type and arrangement of particles in one atom of each of the first 20 elements.

Sample questions

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1 State whether the following are elements, mixtures or compounds.

a silver [1]
b bronze [1]
c seawater [1]
d water [1]
e bauxite [1]
f aluminium oxide [1]

Student's answers

a element
b mixture
c mixture
d compound
e mixture
f compound

Teacher's comments

All the student's answers are correct.

a Silver is an element with atomic number 47. If you are not sure if a substance is an element, you should remember that the Periodic Table only contains elements.

b Bronze is an alloy (see Chapter 10) of copper, tin and other metals in variable proportions and, as such, is a mixture.

c Seawater is water (which is a compound) with many substances, in variable proportions, dissolved in it. Therefore, seawater is a mixture.

d Although water is found in many forms (such as tap water, seawater and distilled water), in Chemistry, *water* refers to the pure substance. Water has the formula H_2O. Any substance with a formula that shows more than one element is a compound.

e Bauxite is a metallic ore from which aluminium is extracted (see Chapter 5). The word *ore* refers to an impure substance and so bauxite is a mixture.

f Bauxite contains the compound aluminium oxide together with impurities. Although bauxite is a mixture, aluminium oxide has the formula Al_2O_3 and, therefore, is a compound.

2 a Complete the table below, showing the charge on each particle.

Particle	Number of protons	Number of electrons	Number of neutrons	Charge on particle
A	20	18	18	
B	17	17	20	
C	9	9	10	
D	17	17	18	
E	13	10	14	
F	16	18	19	

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b State the nucleon number of E. [1]
c Give the letters of the two particles that are isotopes. [2]
d State the name of the element that contains particles of F. [1]

Student's answers

a

Particle	Number of protons	Number of electrons	Number of neutrons	Charge on particle
A	20	18	18	2+
B	17	17	20	0
C	9	10	10	1–
D	17	17	18	0
E	13	10	14	3+
F	16	18	19	2–

b 23
c C and E
d potassium

Teacher's comments

a The student's answers are correct.
- Atoms have equal numbers of protons and electrons and so have no charge.
- Positive ions have more protons than electrons.
  Number of positive charges = number of protons – number of electrons
- Negative ions have fewer protons than electrons.
  Number of negative charges = number of electrons – number of protons

b The student added the numbers of protons and electrons together. The nucleon number is the sum of the numbers of protons and neutrons.
c The student decided incorrectly that isotopes contain the same number of electrons. Isotopes are atoms with the same number of protons.
d The student used the number of neutrons instead of the number of protons to identify the element.

Correct answers

a See student's answer.
b 27
c B and D
d sulfur

Exam-style questions

1 State whether the following are elements, mixtures or compounds.
a iron(III) oxide [1]
b hematite (see Chapter 10) [1]
c iron [1]
d stainless steel (see Chapter 10) [1]
e air [1]

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f oxygen [1]
g natural gas (see Chapter 6) [1]
h methane (see Chapter 12) [1]
[Total: 8]

2 State the number of atoms of each element that are present in the formulae of the following compounds.
a H_2SO_4 [1]
b C_2H_5OH [1]
c $CuSO_4.5H_2O$ [1]
d $Mg(NO_3)_2$ [1]
[Total: 4]

3 Balance the following chemical equations. Some of the balancing numbers have been added for you.
a $CaC_2(s) + __ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$ [1]
b $__ KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + __ H_2O(l)$ [1]
c $TiCl_4(l) + __ Na(s) \rightarrow Ti(s) + __ NaCl(s)$ [1]
d $__ KO_2(s) + __ CO_2(g) \rightarrow 2K_2CO_3(g) + __ O_2(g)$ [1]
e $__ Al(s) + 6HCl(aq) \rightarrow __ AlCl_3 + __ H_2(g)$ [1]
[Total: 5]

4 Complete the table below. Use a Periodic Table if required. [Total: 10]

Element	Number of protons in one atom	Atomic number	Number of neutrons in one atom	Nucleon number
Calcium	20	**a**	21	**b**
Copper	**c**	29	**d**	63
e	29	**f**	36	**g**
Zinc	**h**	**i**	35	**j**

5 Draw a labelled diagram to show the atomic structure of an atom of $^{7}_{3}Li$.
Show the particles in the nucleus as well as the electrons. [Total: 3]

6 Boron exists as two isotopes.
a State the meaning of the term *isotopes*. [1]
b Use the data in the table below to calculate the relative atomic mass (A_r) of boron. ALL WORKING OUT MUST BE SHOWN. [3]
[Total: 4]

Relative mass (mass number)	Percentage abundance
10	20
11	80

7 Copper has two isotopes with mass numbers 63 and 65.
The relative atomic mass (A_r) of copper is 63.5.
Calculate the percentage abundance of each isotope. [Total: 5]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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3 Bonding and structure

Key objectives

By the end of this section, you should be able to:

Ionic bonding and structure

- state that an ionic bond is a strong electrostatic attraction between oppositely charged ions
- describe the formation of ionic bonds between elements from Groups I and VII, including use of dot-and-cross diagrams
- explain the melting points, boiling points and electrical conductivity of ionic compounds in terms of structure and bonding

- describe the formation of ionic bonds between ions of metallic and non-metallic elements, including use of dot-and-cross diagrams
- describe the giant lattice structure of ionic compounds as a regular arrangement of alternating positive ions and negative ions
- deduce the formulae of ionic compounds from:
  - the relative numbers of ions present in a model or a diagram
  - the charges on the ions

- use a Roman numeral to indicate the oxidation number of an element in a compound
- identify oxidation and reduction in redox reactions

- define oxidation in terms of loss of electrons and increase in oxidation number
- define reduction in terms of gain of electrons and decrease in oxidation number
- identify redox reactions:
  - as reactions involving gain and loss of electrons
  - by calculating changes in oxidation number
  - by the colour changes involved using acidified aqueous potassium manganate(VII) or aqueous potassium iodide
- identify oxidising and reducing agents in redox reactions

Covalent bonding and structure

- describe a covalent bond in terms of electrons
- describe and use dot-and-cross diagrams to show the formation of covalent bonds in simple molecules, including H_2, Cl_2, H_2O, CH_4, NH_3 and HCl
- explain the melting points and boiling points, and electrical conductivity, of simple molecular structure compounds in terms of structure and bonding
- describe the giant covalent structures of graphite and diamond
- relate the structure and bonding of graphite and diamond to the use of:
  - graphite as a lubricant and as an electrode
  - diamond in cutting tools

- describe and use dot-and-cross diagrams to show the formation of more complex covalent molecules, such as CH_3OH, C_2H_4, O_2, CO_2 and N_2
- describe the giant covalent structure of silicon(IV) oxide (silicon dioxide)
- relate the similar properties of diamond and silicon(IV) oxide to their structures

Metallic bonding and structure

- describe metallic bonding as the electrostatic attraction between the positive ions in a giant metallic lattice and a mobile 'sea' of delocalised electrons
- explain the electrical conductivity, malleability and ductility of metals in terms of structure and bonding

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Key terms

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Term	Definition
Covalent bond	A chemical bond formed by the sharing of one or more pairs of electrons between two atoms.
Delocalised electrons	Electrons that are spread out within a metal structure. The electrons are not attached to any one particular ion.
Giant ionic lattice	A regular arrangement of positive and negative ions held together by the electrostatic forces of attraction between ions.
Intermolecular force	A weak force of attraction between simple molecules.
Ionic bond	A strong electrostatic force of attraction between oppositely charged ions.
Lattice	A regular arrangement (repeating pattern) of atoms, molecules or ions in a solid.
Oxidation	Gain of oxygen OR loss of electrons OR increase in oxidation number.
Oxidising agent	A substance that oxidises another substance and is itself reduced.
Reducing agent	A substance that reduces another substance and is itself oxidised.
Reduction	Loss of oxygen OR gain of electrons OR decrease in oxidation number.

3.1 Ionic bonding

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Ionic bonding occurs in compounds containing metallic elements combined with non-metallic elements.

- Metal atoms (with 1, 2 or 3 electrons in their outer shells) *lose* an electron or electrons, leaving them with a full outer shell (a noble gas electronic configuration). They form positive ions (**cations**).

- Loss of electrons is known as **oxidation**.

- Non-metal atoms (with 5, 6 or 7 electrons in their outer shells) *gain* an electron or electrons, leaving them with a full outer shell. They form negative ions (**anions**).

- Gain of electrons is known as **reduction**.

Sodium chloride

Sodium atoms contain 11 protons and 11 electrons. Chlorine atoms contain 17 protons and 17 electrons. As both contain equal numbers of protons and electrons, both atoms are uncharged.

Figure 3.1 shows an electron moving from the outer shell of a sodium atom to the outer shell of a chlorine atom, leaving both atoms with a full outer shell (see Figure 3.1).

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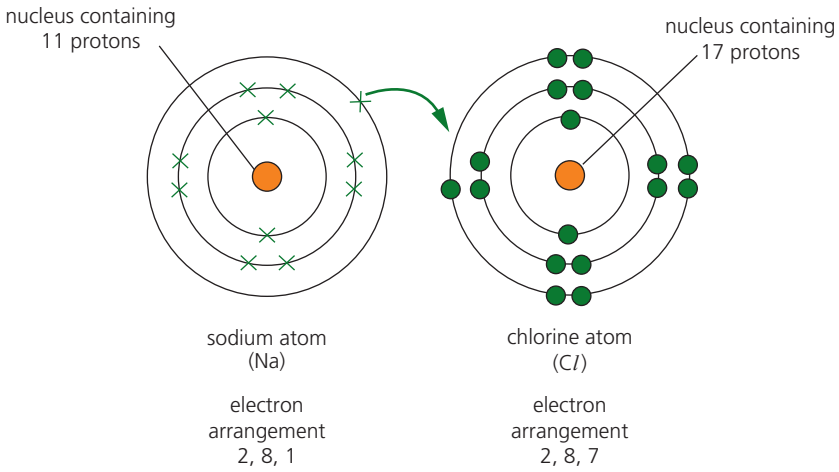
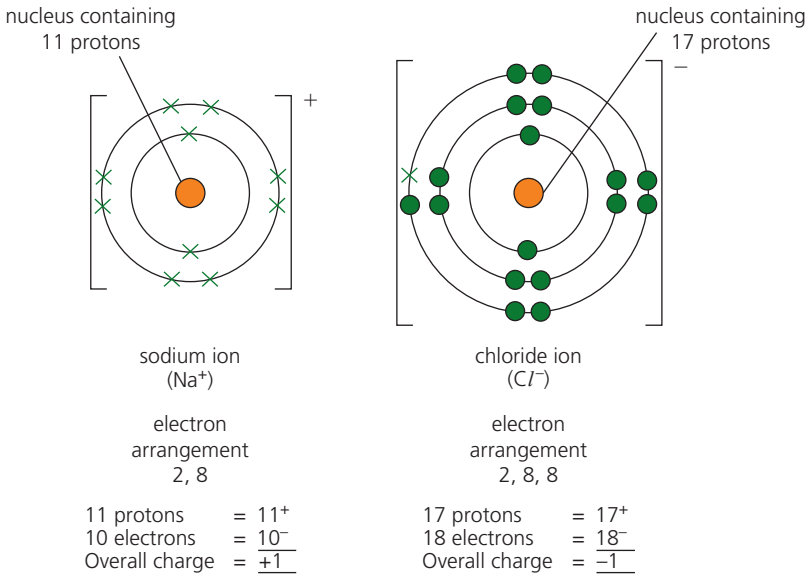


▲ **Figure 3.1 Movement of an electron between a sodium atom and a chlorine atom**

After the transfer of an electron, sodium forms a positive sodium ion and chlorine forms a negative chlor**ide** ion (see Figure 3.2).



▲ **Figure 3.2 The electron arrangements of the resulting sodium and chloride ions**

The sodium ion still has 11 protons but only 10 electrons, therefore it has one positive charge and is written Na^+. The chloride ion still has 17 protons but now has 18 electrons, therefore it has one negative charge and is written Cl^-.

In sodium chloride, the ratio of sodium ions to chloride ions is 1:1. Thus, the formula of sodium chloride is NaCl.

Whenever Group I elements combine with Group VII elements, the ratio of ions is always 1:1.

Magnesium fluoride

In some ionic compounds, atoms combine in different ratios. This happens when the number of electrons lost by one metal atom is not equal to the number of electrons gained by one non-metallic atom.

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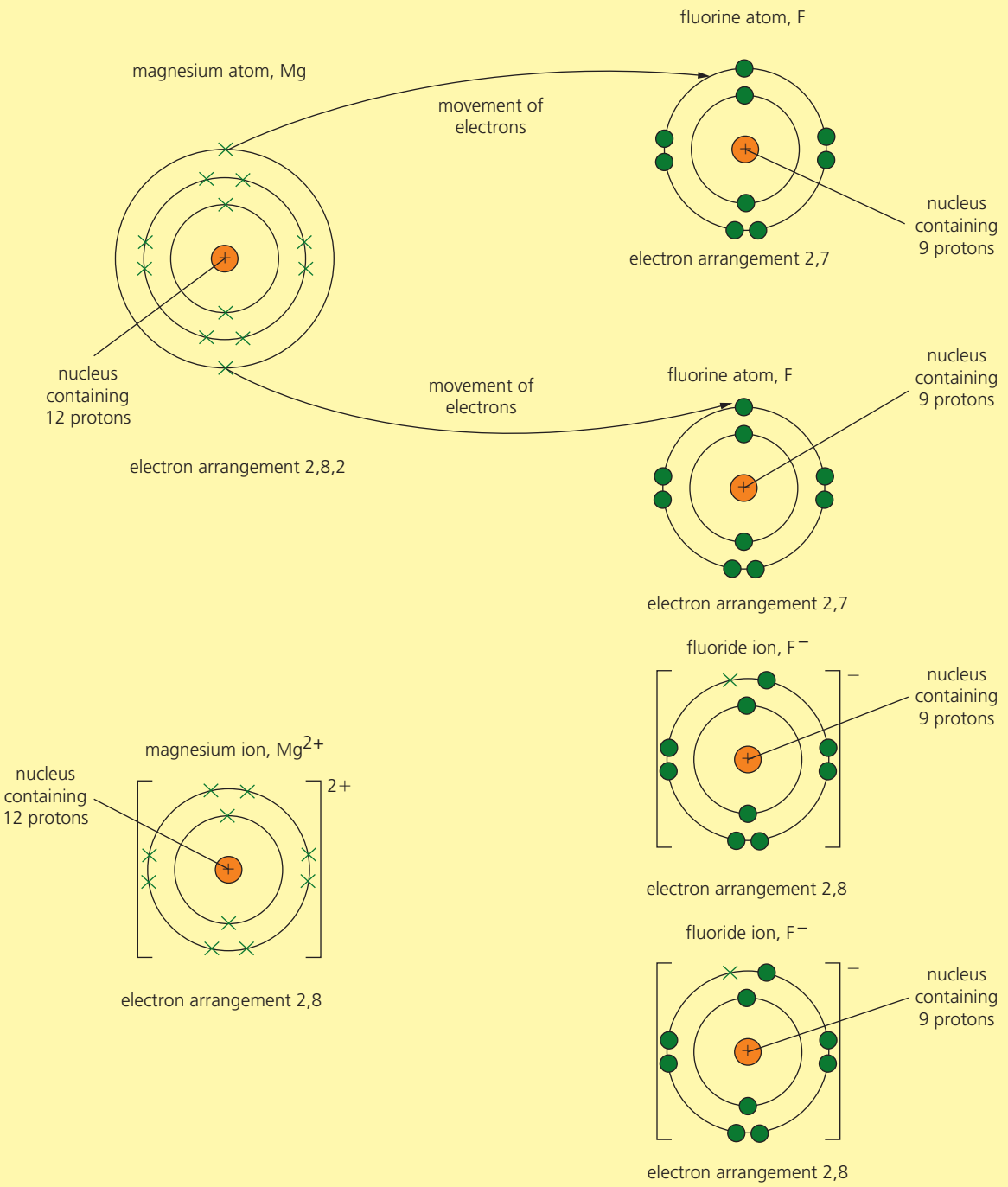

▲ **Figure 3.3 Formation of magnesium fluoride**

As the ratio of magnesium ions to fluoride ions is 1:2, the formula of magnesium fluoride is MgF_2.

The formulae of ionic compounds

The formulae of ionic compounds can be deduced from knowledge of the charges on the ions. The charges on some common ions are shown in Table 3.1.

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Table 3.1 Charges on some common ions

1+	2+	3+	1–	2–	3–
Lithium, Li^+	Magnesium, Mg^{2+}	Aluminium, Al^{3+}	Fluoride, F^-	Oxide, O^{2-}	Nitride, N^{3-}
Sodium, Na^+	Calcium, Ca^{2+}	Iron(III), Fe^{3+}	Chloride, Cl^-	Sulfide, S^{2-}	Phosphate, $\mathbf{PO_4^{3-}}$
Potassium, K^+	Barium, Ba^{2+}		Bromide, Br^-	Carbonate, $\mathbf{CO_3^{2-}}$	
Silver, Ag^+	Zinc, Zn^{2+}		Iodide, I^-	Sulfate, $\mathbf{SO_4^{2-}}$	
Ammonium, $\mathbf{NH_4^+}$	Iron(II), Fe^{2+}		Hydroxide, $\mathbf{OH^-}$	Sulfite, $\mathbf{SO_3^{2-}}$	
	Copper(II), Cu^{2+}		Nitrite, $\mathbf{NO_2^-}$		
	Lead, Pb^{2+}		Nitrate, $\mathbf{NO_3^-}$		

Bold type denotes polyatomic ions. These are ions which have more than one capital letter in the formula (see Rule 4 below).

Skills

Working out formulae

The most important thing to remember is that all compounds have no overall charge. Therefore, in the case of ionic compounds, the number of positive charges is equal to the number of negative charges.

To work out the formula of a compound you should follow these rules:

1. Write down the formulae of the positive and negative ions.
2. Count the number of positive charges and the number of negative charges.
3. If the charges are *not* equal, add more positive ions, more negative ions or both until the charges are equal.
4. If more than one of a polyatomic ion is required, the whole formula of the ion must go in brackets and the number of ions goes outside the brackets as a subscript, e.g. $(NO_3)_2$.

Worked example 1

Deduce the formula of sodium carbonate.

Answer

1 Ions: Na^+ CO_3^{2-}
2 Charges: Na^{+} CO_3^{2-}
1+ 2–
3 Add 1 extra Na^+ to make the charges equal.
Na^{+} CO_3^{2-}
Na^{+}
2+ 2–
4 There is only one CO_3^{2-}, so no brackets are needed.
5 Formula: Na_2CO_3

Worked example 2

Deduce the formula of magnesium hydroxide.

Answer

1 Ions: Mg^{2+} OH^-
2 Charges: Mg^{2+} OH^{-}
2+ 1–
3 Add 1 extra OH^- to make the charges equal.
Mg^{2+} OH^{-}
OH^{-}
2+ 2–
4 OH^- has two capital letters. Since 2 OH^- ions are needed, OH goes in brackets with the 2 outside as a subscript.
5 Formula: $Mg(OH)_2$

Worked example 3

Deduce the formula of aluminium oxide.

Answer

1 Ions: Al^{3+} O^{2-}
2 Charges: Al^{3+} O^{2-}
3+ 2–
3 Al^{3+} O^{2-}
Al^{3+} O^{2-}
O^{2-}
6+ 6–
4 There are no polyatomic ions, so no brackets are needed.
5 Formula: Al_2O_3

Worked example 4

Deduce the formula of iron(III) sulfate.

Answer

1 Ions: Fe^{3+} SO_4^{2-}
2 Charges: Fe^{3+} SO_4^{2-}
3+ 2–
3 Fe^{3+} SO_4^{2-}
Fe^{3+} SO_4^{2-}
SO_4^{2-}
6+ 6–

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4 As SO_4^{2-}, has two capital letters, it goes in brackets with the 3 outside as a subscript.
5 Formula: $Fe_2(SO_4)_3$

These incorrect formulae for iron(III) sulfate show mistakes that are easy to make:

- $FeSO_4$: The number of charges has not been made equal. (This is the correct formula of iron(II) sulfate.)
- $Fe_2(SO)_3$: The 4 is left out of the formula of sulfate.
- $(Fe)_2(SO_4)_3$: No brackets are needed around Fe as it only has one capital letter.
- $Fe_2(SO)_4$: The 4 is left out of the formula of sulfate and placed incorrectly outside the brackets.

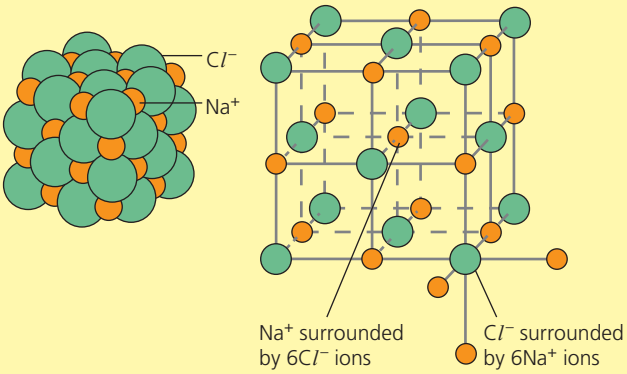
Structure of ionic substances

The ions in sodium chloride, and all other solid substances made from ions, are arranged in a **giant ionic lattice** (see Figure 3.4). A **lattice** is a regular arrangement of particles forming a repeated pattern.

The sodium chloride lattice is held together by strong electrostatic forces of attraction between oppositely charged sodium ions and chloride ions. These forces of attraction are called **ionic bonds**.

Revision activity

Choose one positive ion and one negative ion from Table 3.1. Work out the formula of the compound they form. Then try another pair. How many can you make?



▲ **Figure 3.4 The sodium chloride lattice**

Properties of ionic substances

Ionic substances have:

- high melting points and boiling points
- good electrical conductivity when aqueous or molten
- poor electrical conductivity when solid

Explaining the properties of ionic substances

- High melting points and boiling points are due to strong electrostatic forces of attraction between oppositely charged ions.
- Substances only conduct electricity if they contain moving charged particles (see Chapter 5). Ionic compounds contain positively and negatively charged ions.
  - In the solid state, the ions are held together by strong electrostatic forces. Therefore, the ions are not moving and this explains why solid ionic compounds are poor conductors.
  - When ionic compounds are dissolved in water (aqueous) or heated until they become liquid (molten), the ions are moving.

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Oxidation

Oxidation numbers

Roman numerals in the names of compounds indicate the oxidation numbers of the elements. For example, in this reaction, iron(II) oxide becomes iron(III) oxide:

iron(II) oxide + oxygen → iron(III) oxide

$$4FeO(s) + O_2(g) \rightarrow 2Fe_2O_3(s)$$

The oxidation number of iron changes from +2 to +3. An increase in oxidation number is another definition of oxidation. Therefore, in this reaction, iron(II) oxide is oxidised to iron(III) oxide and oxygen is the **oxidising agent**.

Skills

Working out oxidation numbers

The rules for determining oxidation numbers are as follows:

1. The oxidation number of an uncombined element is 0.
2. In a compound, some elements always have the same oxidation numbers:
   - Group I elements are always +1
   - Hydrogen is always +1 *
   - Group II elements are always +2
   - Aluminium is always +3
   - Fluorine is always –1
   - Oxygen is always –2 *

   * *This is true for the compounds you will meet at IGCSE.*
3. The oxidation number of a monatomic ion is equal to the charge on the ion. For example:
   - in Cu^+ the oxidation number of copper is +1
   - in S^{2-} the oxidation number of sulfur is –2
4. The sum of the oxidation numbers in a polyatomic ion is equal to the charge on the ion. For example, the total oxidation number of an SO_4^{2-} ion is –2.
5. The sum of the oxidation numbers in a compound is 0.

Worked example

Give the oxidation number of nitrogen, N, in each of the following:

a N_2 **b** N_2O_4 **c** NO_3^-

Answers

a N_2 is an element.
The oxidation number of an uncombined element is 0 (Rule 1).
The oxidation number of nitrogen, N, in N_2 is 0.

b N_2O_4 is a compound.
The sum of the oxidation numbers in a compound is 0 (Rule 5).
Oxygen, O, has an oxidation number –2 (list in Rule 2).
Writing this mathematically:

$$N_2O_4 = 0 \text{ and } N_2 + (-2 \times 4) = 0$$
$$N_2 - 8 = 0$$
$$N_2 = +8$$
$$N = +8 \div 2 = +4$$

The oxidation number of nitrogen, N, in N_2O_4 is +4.

c NO_3^- is a polyatomic ion.
The sum of the oxidation numbers is –1 (Rule 4).
Oxygen, O, has an oxidation number –2 (list in Rule 2).

$$N + (-2 \times 3) = -1$$
$$N - 6 = -1$$
$$N = -1 + 6$$
$$N = +5$$

Therefore, the oxidation number of nitrogen, N, in NO_3^- is +5.

Further definitions of redox reactions

Gain of oxygen and/or loss of hydrogen are very limited definitions of oxidation because many redox reactions do not involve oxygen or hydrogen. For example, in the following reaction, iron(II) chloride is oxidised to iron(III) chloride using chlorine as the oxidising agent:

iron(II) chloride + chlorine → iron(III) chloride

$$2FeCl_2(s) + Cl_2(g) \rightarrow 2FeCl_3(s)$$

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The ionic equation for this reaction (see Chapter 8) is:

$$2Fe^{2+}(s) + 3Cl_2(g) \rightarrow 2Fe^{3+}(s) + 6Cl^-(s)$$

This can be broken down into two ionic half-equations (see Chapter 5):

$$\text{Oxidation: } Fe^{2+}(s) \rightarrow Fe^{3+}(s) + e^-$$

$$\text{Reduction: } Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$$

The oxidation number of Fe changes from +2 to +3. Therefore, we can define oxidation as an *increase in oxidation number*.

Fe^{2+} is oxidised to Fe^{3+} by loss of electrons. Cl_2 is the oxidising agent. Therefore, we can also say:

- Oxidation is electron loss.
- Oxidising agents are electron acceptors.

The oxidation number of *Cl* changes from 0 to –1. Therefore, we can define reduction as a *decrease in oxidation number*.

Cl_2 is reduced to $2Cl^-$ by gain of electrons. Fe^{2+} is the **reducing agent**. Therefore, we can also say:

- Reduction is electron gain.
- Reducing agents are electron donors.

In any redox reaction, electrons are transferred from the reducing agent to the oxidising agent. The reducing agent is oxidised and the oxidising agent is reduced.

Testing for oxidising and reducing agents

Aqueous potassium manganate(VII) is an oxidising agent which can be used to test for the presence of reducing agents. When a reducing agent is added, the aqueous potassium manganate(VII) changes from purple to colourless.

Aqueous potassium iodide is a reducing agent which can be used to test for the presence of oxidising agents. When an oxidising agent is added, the aqueous potassium iodide changes from colourless to brown.

3.2 Covalent bonds

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Covalent bonding occurs in elements and compounds containing non-metallic elements only.

Covalent bonds are formed when pairs of electrons are shared. One shared pair of electrons is a single covalent bond.

Double bonds (two shared pairs of electrons) and triple bonds (three shared pairs of electrons) also exist.

Atoms which form a covalent bond join together to form uncharged molecules. All the atoms in the molecule have a full outer shell of electrons (noble gas electronic configurations) because of the shared pairs.

Simple molecules

Figure 3.5 shows dot-and-cross diagrams for some **simple molecules** containing only single bonds. Only the outer electron shells are shown.

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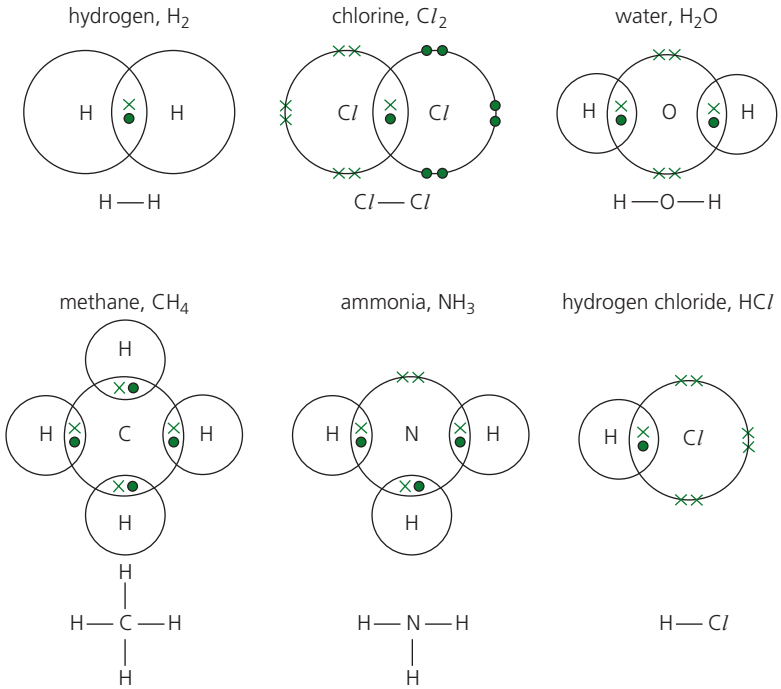
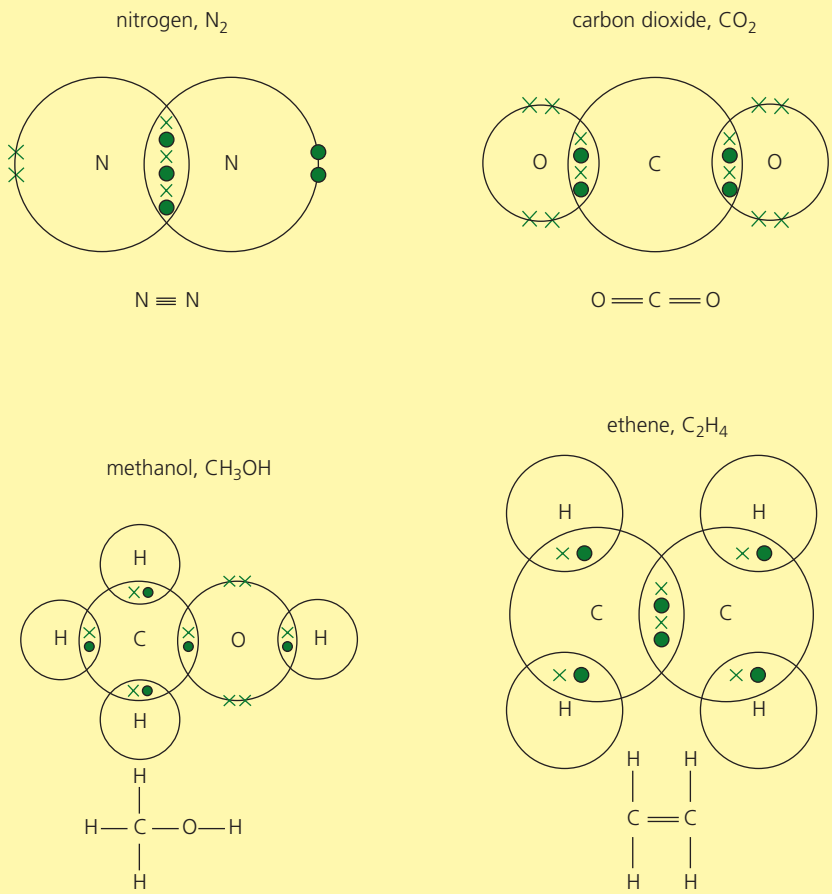


▲ **Figure 3.5 Dot-and-cross diagrams for simple molecules with single bonds**

Figure 3.6 shows dot-and-cross diagrams for more simple molecules, some of which contain double and triple bonds. Again, only the outer electron shells are shown.



▲ **Figure 3.6 Dot-and-cross diagrams for simple molecules with multiple bonds**

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Properties of simple molecular substances

Sulfur is an example of a substance with a simple molecular structure. It has strong covalent bonds between the atoms within the molecules (strong intramolecular bonds) but weak **intermolecular forces** of attraction between the molecules.

Properties of simple molecular substances are shown in Table 3.2.

▼ **Table 3.2 Properties of substances made of simple molecules**

Property	Reason
Low melting points and boiling points	Weak intermolecular attraction between molecules
Poor electrical conductivity	Made of uncharged molecules

Intermolecular forces are only present between simple molecules. They do *not* exist in ionic substances, substances with giant covalent structures or metals.

Substances containing simple molecules are either solids with low melting points (such as iodine), liquids (such as water) or gases (such as carbon dioxide). In exam questions which ask why these have low melting points and boiling points, a common explanation is that covalent bonds are weak. This is a bad error. Remember that:

- all covalent bonds are strong
- covalent bonds do not break when simple molecular substances melt
- only the weak intermolecular forces break

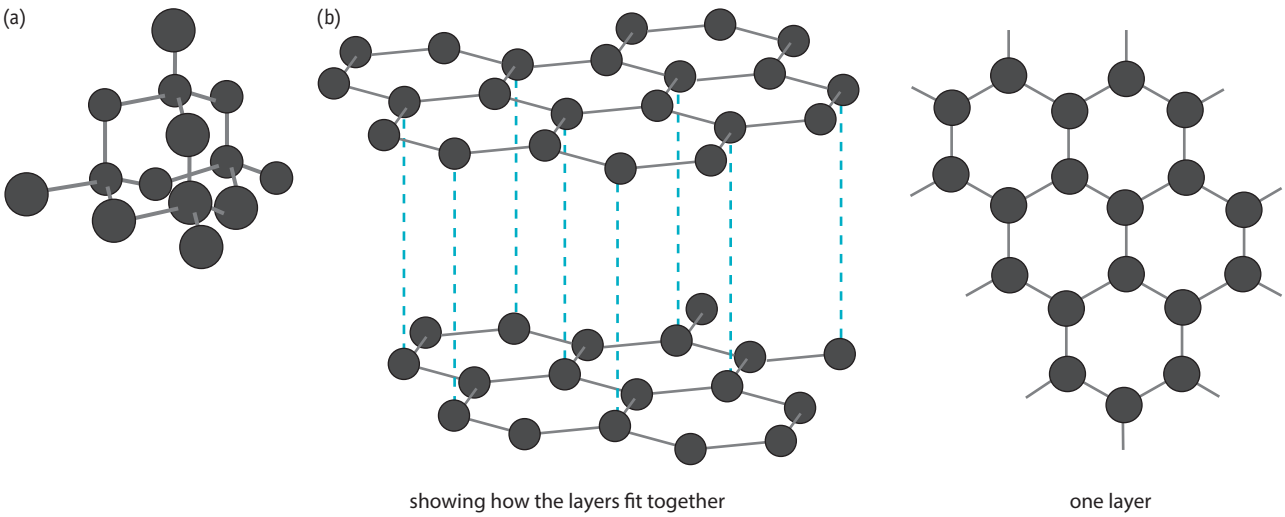
Giant covalent structures

Diamond and graphite

Diamond is an example of a giant covalent structure. It is held together by strong covalent bonds between carbon atoms.

Graphite is another example of a substance with a giant covalent structure.

Differences in the structure and bonding of diamond and graphite lead to their different properties and uses (Table 3.3).



▲ **Figure 3.7 Structures of (a) diamond and (b) graphite**

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▼ **Table 3.3 Differences in structure and bonding of diamond and graphite**

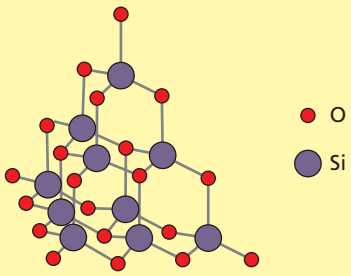
Property	Diamond	Graphite
Number of other carbon atoms covalently bonded to each carbon atom	4	3
Arrangement of atoms	Tetrahedral	In layers – each layer is made of interlocking rings containing six carbon atoms
Bonding	All bonds are covalent	Covalent bonds between atoms within the layers Weak intermolecular forces between the layers
Hardness	Hard because all bonds are strong and directional	Soft because weak intermolecular forces between the layers allow them to slide over each other
Mobile electrons	None – all the outer shell electrons are used in bonding	One electron from each atom is in the spaces between the layers and is mobile
Conduction of electricity	Insulator because there are no mobile electrons	Good conductor due to mobile electrons between layers
Uses	In cutting tools due to hardness and strength	As a lubricant because layers can slide As a conductor in motors

Silicon(IV) oxide, SiO_2

Silicon(IV) oxide, SiO_2, has a **giant covalent structure**.

Each silicon atom is covalently bonded to four oxygen atoms (Figure 3.8). The bonds are directed tetrahedrally. Each oxygen atom is covalently bonded to two silicon atoms.

All the bonds in silicon(IV) oxide are strong covalent bonds. There are no mobile electrons. As a result, silicon(IV) oxide is strong, hard, has high melting and boiling points and is an electrical insulator. These properties are like those of diamond, which has a very similar structure and the same type of bonding.



▲ **Figure 3.8 The structure of silicon(IV) oxide**

3.3 Metallic bonding

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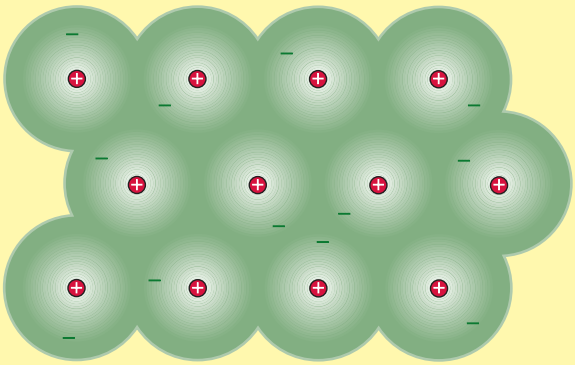
All metallic elements have **giant metallic structures**.

They contain positive ions in a giant metallic lattice and a mobile 'sea' of **delocalised electrons**. Metals are held together by the strong forces of attraction between positive ions and the sea of mobile electrons. These forces are known as **metallic bonds**.



▲ **Figure 3.9 Metallic bonding**

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In exam questions which ask for the meaning of the term *metallic bonding*, remember to describe the strong electrostatic attraction between positive ions and the mobile sea of electrons, not just the lattice structure.

The properties of metals, and an explanation for the properties, are shown in Table 3.4.

▼ **Table 3.4 Explaining the properties of metals**

Property	Explanation
Good electrical conductivity	Delocalised electrons are able to move
Malleability and ductility	Layers of positive ions can slide over one another when a force is applied

Notice that conduction of electricity in metals is *not* caused by the movement of ions – it is due to the movement of electrons only.

Revision activity

Choose three pure substances with different properties. For each one, decide if it contains simple molecules, giant covalent molecules or ions. Discuss the reasons for your decisions with a friend.

Sample questions

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1 Use the table below to answer the questions that follow. Take room temperature as 25°C.

	Melting point/°C	Boiling point/°C	Conducts electricity when solid?	Conducts electricity when molten?
A	–40	35	No	No
B	50	150	No	No
C	801	1500	No	Yes
D	1500	2500	Yes	Yes
E	2500	5000	No	No
F	–75	–35	No	No

a Which substance or substances are solid at room temperature? [4]
b Which substance or substances are liquid at room temperature? [1]
c Which substance or substances are gaseous at room temperature? [1]
d Which substance could have a giant metallic structure? [1]
e Which substance has a giant ionic lattice? [1]
f Which substance has a giant covalent structure? [1]

Student's answers

a	A, B, C, D, E	*d*	D
b	*none*	*e*	E
c	F	*f*	D

Teacher's comments

a The student included A. This was possibly because they ignored the negative sign.
b As the student thought A was a solid, they did not identify it as a liquid.
c The student's answer is correct.
d The student's answer is correct.
e The student correctly looked for a high melting point, but ignored the fact that substances with a giant ionic lattice only conduct electricity when molten.
f The student correctly looked for a high melting point, but ignored the fact that substances with a giant covalent structure do not conduct electricity. (Graphite does conduct, but it has a much higher melting point.)

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Correct answers

a B, C, D, E
b A
c F
d D
e C
f E

2 Carbon dioxide, CO_2, is a simple molecular substance.
Use this information to explain why:
a Carbon dioxide has a low melting point.
b Carbon dioxide is a non-conductor of electricity.

Student's answers

a *Carbon dioxide has weak covalent bonds.*
b *The ions in carbon dioxide cannot move.*

Teacher's comments

a It is a common error to state that covalent bonds are weak.
b Carbon dioxide is made of molecules as opposed to ions.

Correct answers

a The forces of attraction between molecules (the intermolecular forces) are weak.
b Carbon dioxide contains uncharged molecules.

3 State how the structure and bonding of a metallic element, such as iron, cause the metal to have a high melting point.

Student's answer

The structure of metals is referred to as a metallic lattice. Metals contain metal ions and free electrons. The strong attraction results in a high melting point.

Teacher's comments

The student's answer is incomplete.

- The structure of metals is referred to as a giant metallic lattice.
- The ions in metals should be referred to as positive ions (cations) and the electrons should be referred to as a 'sea' of delocalised electrons. *Free* as an adjective to describe electrons (or ions) should never be used in chemistry exams.
- The strong attraction should be referred to as strong electrostatic attraction between positive ions and the sea of delocalised electrons.

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Correct answer

The main points in the answer should be:

- The structure of metals is called a giant metallic lattice.
- The lattice consists of positive ions surrounded by a sea of delocalised electrons.
- The electrostatic attraction between the positive ions and the sea of delocalised electrons is very strong.
- A large amount of energy is required to break this electrostatic attraction, which is why most metals have a high melting point.

Exam-style questions

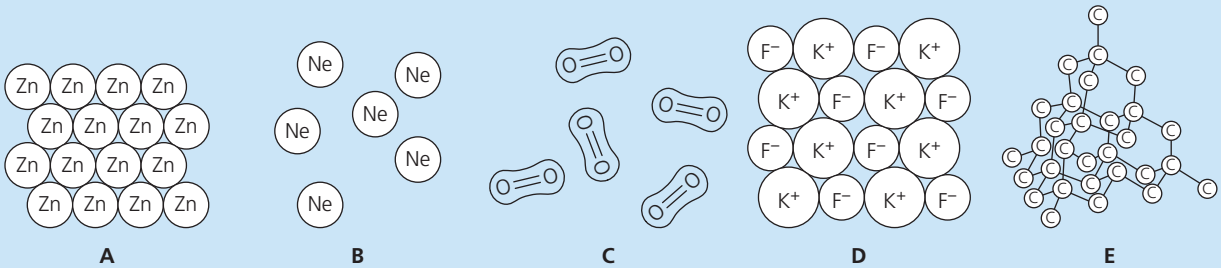
1 Complete the table below. [Total: 5]

Particle	Number of protons	Number of electrons	Electronic configuration	Charge on particle
A	20	18		
B	9	10	2,8	
C		10	2,8	0
D	8	10	2,8	

2 Complete the table below. [Total: 18]

	Giant ionic structures	Giant covalent structures	Simple molecules
Example			
Type of particle present			
Type of bonding between particles			
Melting point and boiling point			
Electrical conductivity when solid			
Electrical conductivity when aqueous		Insoluble in water	
Malleability and ductility		Not malleable or ductile	Not malleable or ductile

3 The diagram below shows the structures of five different substances.



Use the letters **A**, **B**, **C**, **D** and **E** to answer the questions below.
Each letter may be used once, more than once or not at all.
Give the letter that shows:

a atoms with a full outer shell of electrons [1]
b a giant covalent structure [1]
c particles that are formed by loss and gain of electrons [1]
d a substance that conducts electricity when solid [1]

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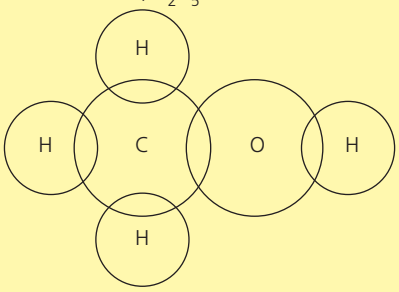
e a substance that only conducts electricity when molten or dissolved in water [1]
f the substance with the lowest melting point [1]
[Total: 6]

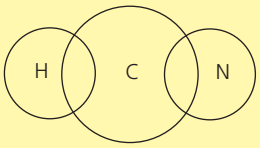
4 Draw the dot-and-cross diagrams to show the electronic configurations (outer shells only) of the following molecules.
a hydrogen fluoride, HF [1]
b fluorine, F_2 [1]
c silicon tetrachloride, $SiCl_4$ (the atoms are arranged in the same way as in methane, CH_4) [1]
d hydrogen sulfide, H_2S (the atoms are arranged in the same way as in water, H_2O) [1]
[Total: 4]

5 Complete the dot-and-cross diagrams below to show the electronic configurations (outer shells only) of the following molecules. [Total: 4]

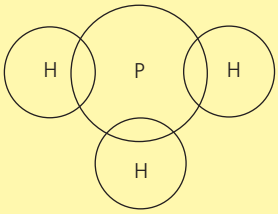
a ethanol, C_2H_5OH



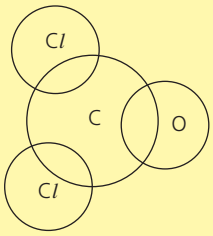
b hydrogen cyanide, HCN



c phosphine, PH_3



d carbonyl chloride, $COCl_2$



6 Deduce the formulae of the following ionic compounds.

a calcium hydroxide
b magnesium chloride
c ammonium phosphate
d lithium sulfide
e lead(II) nitrate
f calcium carbonate
g aluminium nitrate
h potassium sulfite
i zinc sulfate
j ammonium sulfate [Total: 10]

7 The ionic equation (see Chapter 8) for the redox reaction between magnesium and copper(II) sulfate solution is:

$$Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$$

a Write two ionic half-equations representing oxidation and reduction in the reaction. [2]
b Write down the oxidation numbers of magnesium and copper in the following.
i Mg [1] **iii** Cu [1]
ii Mg^{2+} [1] **iv** Cu^{2+} [1]
c i Write down the formula of the substance which is oxidised in the reaction. [1]
ii Explain your answer in terms of:
– electron transfer [1]
– oxidation number [1]
d i Write down the formula of the substance which is reduced in the reaction. [1]
ii Explain your answer in terms of:
– electron transfer [1]
– oxidation number [1]
[Total: 12]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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4 Stoichiometry – chemical equations

Key objectives

By the end of this section, you should be able to:

- define the molecular formula of a compound
- define and calculate relative molecular mass, M_r
- define and calculate relative formula mass, M_r, for ionic compounds
- calculate reacting masses in simple proportions
- state that concentration of solutions can be measured in g/dm^3 or mol/dm^3

- define the empirical formula of a compound
- define the mole and the Avogadro constant
- calculate:
  - amount of substance
  - mass
  - molar mass
  - number of particles using the value of the Avogadro constant
  - volume of gas at r.t.p.
  - volume of solution
  - concentration of solution expressed as g/dm^3 or mol/dm^3, including conversion between cm^3 and dm^3
  - percentage yield
  - percentage composition by mass
  - percentage purity
  - reacting masses to determine which reactant is limiting

Key terms

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Term	Definition
Avogadro constant	6.02×10^{23} The number of atoms, ions or molecules in one mole of a substance.
Empirical formula	A formula showing the simplest whole number ratio of atoms or ions present in a compound.
Molar mass	The mass of one mole of a compound. It has units of g/mol.
Mole	The amount of substance which contains 6.02×10^{23} atoms, ions or molecules.
Molecular formula	A formula showing the actual number and type of different atoms of each element present in one molecule of a compound.
Relative formula mass, M_r	The sum of the relative atomic masses of those elements shown in the formula of any substance.
Relative molecular mass, M_r	The sum of the relative atomic masses of those elements shown in the formula of a molecular substance.

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4.1 Relative atomic mass

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The term *relative atomic mass* was introduced in Section 2.4. If we define a similar term for molecules, we can use the concept to find out more about how substances react.

Relative molecular mass

You can calculate **relative molecular mass**, M_r (also known as **relative formula mass**, M_r, for ionic compounds) if you know the formula of the compound and the relative atomic mass of each element in it.

Skills

Calculating M_r

A compound, iron(II) sulfide, has the formula FeS.

The relative atomic masses (A_r) of iron (Fe) and sulfur (S) are 56 and 32, respectively.

Therefore, the M_r of iron(II) sulfide is 56 + 32 = 88.

Worked examples

Calculate M_r for these compounds:

CO_2 N_2O C_4H_{10} $Pb(NO_3)_2$ $Al_2(SO_4)_3$

Use the following values of A_r:

H = 1, C = 12, N = 14, O = 16, Al = 27, S = 32, Pb = 207

Answers

$CO_2 = 12 + (16 \times 2) = 44$

$N_2O = (14 \times 2) + 16 = 44$

$C_4H_{10} = (12 \times 4) + (1 \times 10) = 58$

Multiplying out the brackets for $Pb(NO_3)_2$ gives:

$PbN_2O_6 = 207 + (14 \times 2) + (16 \times 6) = 331$

Multiplying out the brackets for $Al_2(SO_4)_3$ gives:

$Al_2(SO_4)_3 = Al_2S_3O_{12} = (27 \times 2) + (32 \times 3) + (16 \times 12) = 342$

Skills

Calculating reacting masses

We can convert relative atomic mass and relative molecular/formula mass to actual mass by adding the mass unit, g. This allows us to work out the masses of the substances needed for or produced in a reaction.

Worked example

Calculate the mass of sodium chloride (NaCl) that can be produced when 4.6 g of sodium (Na) burns in excess chlorine, Cl_2.

The equation for the reaction is:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Answer

From a Periodic Table, A_r for Na = 23 and A_r for Cl = 35.5.

So, M_r of NaCl = 23 + 35.5 = 58.5.

Excess chlorine means that there is more than enough chlorine to react with all the sodium.

The equation shows that 2Na produces 2NaCl.

Therefore, 2Na = 2 × 23 = 46 g of Na produces 2NaCl = 2 × 58.5 = 117 g of NaCl.

Therefore, 4.6 g of Na produces $\frac{4.6 \times 117}{46} = 11.7$ g of NaCl.

Revision activity

Create an example of your own for each type of calculation. You could look in other chapters for examples of compounds and reactions to use. Work out the answers, too. Then swap your questions with a classmate. Did you get the same answers? If not, work together to find the mistakes.

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4.2 Calculating moles

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A **mole** is a useful way of measuring the amount of a substance.

One mole of any substance contains 6.02×10^{23} particles. This number is known as the **Avogadro constant**, N_A. The relationship can be written as an equation:

$$\text{number of moles} = \frac{\text{number of particles}}{\text{Avogadro constant}}$$

Moles and elements

Some elements (all metals and some non-metals, such as sulfur, S, and carbon, C) exist as atoms.

Other elements, such as hydrogen, H_2, and oxygen, O_2, exist as molecules.

- For any substance which contains only atoms, the mass of one mole is the same as the relative atomic mass, A_r, in grams.
- If the substance is made of molecules, the mass of one mole is the same as the relative molecular mass or relative formula mass, M_r, in grams (see Section 4.3).

Skills

Calculating amount from mass

We can write the information above as an equation:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g/mol)}}$$

Worked example

Calculate the number of moles in 6.0 g of carbon atoms. A_r of C = 12.

Answer

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g/mol)}}$$

$$= \frac{6\text{ g}}{12\text{ g/mol}} = 0.50\text{ moles}$$

Notice that the answer is given to the same number of significant figures as used in the question.

Skills

Calculating number of particles from mass

Since one mole of a substance contains N_A particles:

$$\text{number of particles} = \text{number of moles} \times \text{Avogadro constant}$$

Worked example

Calculate the number of molecules in 3.0 g of hydrogen molecules, H_2.

Answer

M_r of $H_2 = (1 \times 2) = 2$

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g/mol)}}$$

$$= \frac{3.0\text{ g}}{2\text{ g/mol}} = 1.5$$

$$\text{number of particles} = \text{number of moles} \times \text{Avogadro constant}$$

$$= 1.5 \times 6.02 \times 10^{23}$$

$$= 9.03 \times 10^{23}\text{ molecules}$$

4.3 Moles and compounds

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The number of moles in compounds can be also calculated from:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g/mol)}}$$

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The **molar mass** of a compound is the relative molecular mass (for covalent compounds) or relative formula mass (for ionic compounds), M_r, in grams.

Skills

Calculating mass from moles

We can rearrange the equation above for mass:

$$\text{mass (in g)} = \text{number of moles} \times \text{molar mass (in g/mol)}$$

Note that the mass must be in grams (g). If the mass is given in kilograms (kg), it must be multiplied by 1000 and if it is given in tonnes, it must be multiplied by 1 000 000 to convert it into grams.

Worked example

Calculate the mass of 1.5 moles of carbon dioxide molecules, CO_2.

Answer

M_r of $CO_2 = 12 + (16 \times 2) = 44$

molar mass = 44 g/mol

mass of 1.5 moles = 1.5 × 44 g/mol = 66 g

You should also be able to calculate the number of moles from:

- mass
- number of particles
- volume of a gas
- volume and concentration of a solution

Moles and gases

The volume of one mole of any gas at room temperature and pressure (r.t.p.) is 24 dm^3 = 24 000 cm^3. This applies to both elements and compounds.

Skills

Mole calculations using volume of a gas

Writing the above information as an equation:

$$\text{number of moles of a gas} = \frac{\text{volume of the gas (in dm}^3\text{ at r.t.p.)}}{24\,\text{dm}^3}$$

Rearranging for volume of a gas:

$$\text{volume of the gas (in dm}^3\text{ at r.t.p.)} = \text{number of moles of a gas} \times 24\,\text{dm}^3$$

The volume of gas and the volume of one mole of gas can be in either cm^3 or dm^3, but you must use the same units for *both*, not one of each.

Worked example

Calculate the number of moles in 120 cm^3 of carbon dioxide gas, CO_2, at r.t.p.

Answer

$$\text{number of moles of a gas} = \frac{\text{volume of the gas (in cm}^3\text{ at r.t.p.)}}{24\,000\,\text{cm}^3}$$

$$\text{number of moles of CO}_2 = \frac{120\ \text{cm}^3}{24\ 000\ \text{cm}^3}$$

$$= 5.00 \times 10^{-3}\ \text{moles}$$

The fact that the gas is carbon dioxide is irrelevant to this question. One mole of *any* gas occupies 24 dm^3 (or 24 000 cm^3) at r.t.p., so the answer would be the same for any gas.

Moles and solutions

The **concentration** of a solution can be expressed in grams per cubic decimetre (g/dm^3), but chemists usually use moles per cubic decimetre (mol/dm^3).

$$\text{number of moles} = \text{volume (in dm}^3) \times \text{concentration (in mol/dm}^3)$$

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Rearranging for volume:

$$\text{volume (in dm}^3) = \frac{\text{number of moles}}{\text{concentration (in mol/dm}^3)}$$

Rearranging for concentration:

$$\text{concentration (in mol/dm}^3) = \frac{\text{number of moles}}{\text{volume (in dm}^3)}$$

Skills

Calculations using concentrations

Correct units are very important in these equations. As solutions are often measured using burettes and pipettes, which are graduated in cm^3, the equations below may be more useful:

$$\text{number of moles} = \frac{\text{volume (in cm}^3) \times \text{concentration (in mol/dm}^3)}{1000}$$

$$\text{volume (in cm}^3) = \frac{\text{number of moles} \times 1000}{\text{concentration (in mol/dm}^3)}$$

$$\text{concentration (in mol/dm}^3) = \frac{\text{number of moles} \times 1000}{\text{volume (in cm}^3)}$$

It is a good idea to remember that the expressions for volume (cm^3) and concentration (mol/dm^3) have moles × 1000 on the top line.

Worked example 1

Calculate the number of moles in 20.0 cm^3 of a solution of aqueous sodium hydroxide, NaOH, of concentration 0.200 mol/dm^3.

Answer

$$\text{number of moles} = \frac{\text{volume (in cm}^3) \times \text{concentration (in mol/dm}^3)}{1000}$$

$$= \frac{20.0\ \text{cm}^3 \times 0.200\ \text{mol/dm}^3}{1000}$$

$$= 4.00 \times 10^{-3}\ \text{moles}$$

The calculation is the same for any solution. The fact that the solution is sodium hydroxide is irrelevant to this question.

Worked example 2

Calculate the volume in cm^3 of 0.250 mol/dm^3 dilute sulfuric acid, $H_2SO_4(aq)$, that contains 0.00500 moles.

Answer

$$\text{volume (in cm}^3) = \frac{\text{number of moles} \times 1000}{\text{concentration (in mol/dm}^3)}$$

$$= \frac{0.00500\ \text{moles} \times 1000}{0.250\ \text{mol/dm}^3}$$

$$= 20.0\ \text{cm}^3$$

The calculation is the same for any solution. The fact that the solution is dilute sulfuric acid is irrelevant to this question.

4.4 Calculating formulae

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Empirical formulae

The **empirical formula** is the formula showing the smallest whole number ratio of atoms or ions present in a compound.

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Skills

Finding empirical formulae

The empirical formula of a compound can be calculated from:

- the masses of the elements that combine to make the compound
- the percentage of each element (by mass) in the compound.

Worked example 1

A compound is made of manganese and oxygen only.

The compound contains 0.33 g of manganese, Mn, and 0.144 g of oxygen, O.

Calculate the empirical formula of the compound.

Relative atomic masses, A_r: Mn = 55, O = 16

Answer

Use the following steps:

1 Find the number of moles of each element using:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g/mol)}}$$

moles of manganese atoms,
Mn $= 0.33 \div 55 = 0.006$

moles of oxygen atoms,
O $= 0.144 \div 16 = 0.009$

If the calculation involves an element that forms diatomic molecules, make sure you use A_r not M_r, so here we use O = 16 not O = 32.

2 Divide the number of moles of each element by the smallest number:

Mn: $0.006 \div 0.006 = 1$

O: $0.009 \div 0.006 = 1.5$

The empirical formula can only have whole numbers of each atom. To say that 1.5 is approximately 1 (giving a formula of MnO) or that 1.5 is approximately 2 (giving a formula of MnO_2) is incorrect. You can only make an approximation like this if the result of the division is very close to a whole number (i.e. 0.1 away or less).

3 If any of the numbers is *not* a whole number, multiply both numbers by a whole number greater than 1 until both give whole numbers.

Mn: $1 \times 2 = 2$

O: $1.5 \times 2 = 3$

Both are now whole numbers, so there is no need to try another number.

4 These numbers tell you the subscripts to use for each element, so you can write down the empirical formula: Mn_2O_3.

Worked example 2

A compound has the following percentage composition by mass: 26.7% carbon, 2.2% hydrogen, 71.1% oxygen.

Deduce the empirical formula of the compound.

Relative atomic masses A_r: C = 12, H = 1, O = 16

Answer

The figures for percentage composition by mass mean that 100 g of the compound contains 26.7 g of carbon, 2.2 g of hydrogen and 71.1 g of oxygen.

Therefore, we can proceed as before.

1 Find the number of moles of each element:

moles of carbon atoms,
C $= 26.7 \div 12 = 2.225$

moles of hydrogen atoms,
H $= 2.2 \div 1 = 2.2$

moles of oxygen atoms,
O $= 71.1 \div 16 = 4.44375$

2 Divide each of the above by the smallest number:

C: $2.225 \div 2.2 = 1$

H: $2.2 \div 2.2 = 1$

O: $4.44375 \div 2.2 = 2$

3 All of these are whole numbers, so we can go straight to Step 4.

4 Write down the empirical formula: CHO_2.

Molecular formulae

The **molecular formula** shows the actual number and type of different atoms of each element present in one molecule of a compound.

Examples of molecular and empirical formulae are shown in Table 4.1.

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▼ **Table 4.1 Molecular and empirical formulae for some compounds**

Compound	Molecular formula	Empirical formula
Butane	C_4H_{10}	C_2H_5
Hydrogen peroxide	H_2O_2	HO
Glucose	$C_6H_{12}O_6$	CH_2O
Benzene	C_6H_6	CH
Methane	CH_4	CH_4

Skills

Determining molecular formulae from empirical formulae

It is possible to determine the molecular formula of a substance from its empirical formula if the relative molecular mass of the substance is also known.

Worked example

The empirical formula of a compound is CH_2 and it has a relative molecular mass of 70. Deduce the molecular formula of the compound.

Answer

The empirical formula is CH_2, so the molecular formula can be expressed as $(CH_2)_n$, where n is a whole number. Therefore:

$$n = \frac{M_r \text{ of the compound}}{M_r \text{ of the empirical formula}}$$

M_r of the compound = 70 and M_r of CH_2 = 12 + (1 × 2) = 14.

Therefore:

$$n = 70 \div 14 = 5$$

and the molecular formula is $(CH_2)_5 = C_5H_{10}$.

Hydrated salts

Hydrated salts are salts containing water of crystallisation (see Section 8.6). An example is hydrated copper(II) sulfate, $CuSO_4.5H_2O$. The formula means that 1 mole of $CuSO_4$ is combined with 5 moles of H_2O.

Skills

Calculating water of crystallisation

The mole concept can be used to calculate the number of moles of water of crystallisation in one mole of a hydrated salt.

Worked example

Hydrated sodium sulfate can be represented as $Na_2SO_4.xH_2O$.

When hydrated sodium sulfate is heated, it gives off water. The water is given off as steam.

$$Na_2SO_4.xH_2O(s) \rightarrow Na_2SO_4(s) + xH_2O(g)$$

The remaining solid is known as anhydrous sodium sulfate, Na_2SO_4.

A student carries out an experiment to determine the value of x.

- The hydrated sodium sulfate is weighed.
- The hydrated salt is heated.
- The remaining solid is weighed.

a Describe what the student should do to make sure that *all* the water is given off.

b 0.805 g of $Na_2SO_4.xH_2O$ is heated until all the water is given off. The mass of Na_2SO_4 remaining is 0.355 g.
Using M_r of Na_2SO_4 = 142 and M_r of H_2O = 18:
- **i** calculate the number of moles of Na_2SO_4 remaining
- **ii** calculate the mass of H_2O given off
- **iii** calculate the number of moles of H_2O given off
- **iv** calculate the value of x.

Answer

a Repeat Steps 2 and 3 until the mass stops decreasing.
A common incorrect answer is to state that water given off should be tested with either anhydrous copper(II) sulfate or anhydrous cobalt(II) chloride. Neither of these will detect very small amounts of water. To ensure that *all* the water is given off, measurement of mass is required.

b A common error in the first part of this calculation is to use the mass of the hydrated salt (0.805 g) instead of the mass of the anhydrous salt (0.355 g).

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i moles of Na_2SO_4 remaining = 0.355 ÷ 142 = 2.5×10^{-3} or 0.0025 moles
ii mass of H_2O given off = 0.805 – 0.355 = 0.450 g
iii moles of H_2O given off = 0.450 ÷ 18 = 2.5×10^{-2} or 0.025 moles
iv x = number of moles of H_2O ÷ number of moles of Na_2SO_4

x = 0.025 ÷ 0.0025 = 10
x must be a whole number. If, when doing an exam question like this, it is not a whole number, check the earlier parts of your calculation.

Always check answers to one part of a question before going on to the next.

4.5 Moles and chemical equations

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Skills

Calculations using moles and chemical equations

Calculations asking about amounts in reactions should be approached in the following order:

1. Calculate any relative molecular masses, M_r, that are required.
2. Calculate the number of moles of the substance where sufficient information is given to do so.
3. Use the mole ratio in the equation to calculate the number of moles of the other substance.
4. Use your answer to Step 3 to calculate what is required, which may be:
   - mass
   - volume of gas
   - volume of solution
   - concentration of solution
   - number of particles

It is extremely important to show all the working out in calculations. If some correct working is shown and the final answer is incorrect, you will still be awarded some of the available marks.

Worked example 1

Calculate the volume of carbon dioxide, CO_2, at room temperature and pressure (r.t.p.) that is produced by heating 2.1 g of sodium hydrogen carbonate, $NaHCO_3$, according to the equation:

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(s) + H_2O(l)$$

A_r of Na = 23, H = 1, C = 12, O = 16

The volume of one mole of any gas is 24 dm^3 at r.t.p.

Answer

M_r of $NaHCO_3 = 23 + 1 + 12 + (16 \times 3) = 84$

moles of $NaHCO_3 = 2.1 \div 84 = 0.025$ moles

From the equation above, 2 moles of $NaHCO_3$ produce 1 mole of CO_2.

So, the mole ratio = 2:1 and therefore:

0.025 moles of $NaHCO_3$: 0.025 ÷ 2 = 0.0125 moles of CO_2

Make sure you use the mole ratio in the equation, taking care not to use it the wrong way round (i.e. 1:2 instead of 2:1). It is possible to score 3 out of 4 marks under these circumstances, depending on how much correct working out you show.

volume of CO_2 = number of moles × volume of 1 mole of gas
= 0.0125 × 24 = 0.3 dm^3

The question asks for the volume of carbon dioxide. It is a very common error to calculate the mass instead. Those who do this can achieve the first 3 marks as long as the working out is clearly shown.

Worked example 2

Calculate the volume of aqueous sodium hydroxide, NaOH(aq), of concentration 0.20 mol/dm^3, which would be required to neutralise exactly 25.0 cm^3 of dilute sulfuric acid, H_2SO_4(aq), of concentration 0.25 mol/dm^3 according to the equation:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

This question asks you to calculate the volume of a solution. In questions like this, watch out for these common mistakes:

- Many students use the value of 24 dm^3 because they confuse the volume of a solution with the volume of a gas.
- When calculating the number of moles of a solution, many students use the equation:

  moles = concentration × volume

  which they may have learnt as $n = c \times v$. This works if the volume is in dm^3, but in this case the volume is in cm^3, which means that a factor of 1000 must be included.

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Answer

There are no masses involved in this question, so no M_r values need to be calculated.

$$\text{moles of } H_2SO_4 = \frac{25.0 \times 0.25}{1000} = 6.25 \times 10^{-3} \text{ moles}$$

The mole ratio in the equation is 1 mole of H_2SO_4 : 2 moles of NaOH.

$$6.25 \times 10^{-3} \times 2 = 0.0125 \text{ moles NaOH}$$

$$\text{volume of NaOH} = \frac{\text{number of moles} \times 1000}{\text{concentration (in mol/dm}^3)}$$

$$= \frac{0.0125 \times 1000}{0.20} = 62.5 \text{ cm}^3$$

Worked example 3

240 dm^3 of nitrogen, $N_2(g)$, reacts with excess hydrogen, $H_2(g)$, according to the equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

a Calculate the volume of ammonia, $NH_3(g)$, produced.

b What volume of hydrogen, $H_2(g)$, would react with the nitrogen?

All volumes are measured at r.t.p.

The volume of one mole of any gas is 24 dm^3 at r.t.p.

Answer

a There are no masses involved in this question, so no M_r values need to be calculated.

moles of nitrogen = 240 ÷ 24 = 10.0

The mole ratio is 1 mole of nitrogen : 2 moles of ammonia, so:

10 moles of nitrogen : 2 × 10 = 20 moles of ammonia

volume of 20 moles of ammonia = 20 × 24 = 480 dm^3

b It is possible to start at Step 3:
The mole ratio is 1 mole of nitrogen : 3 moles of hydrogen, so:

10 moles of nitrogen : 3 × 10 = 30 moles of hydrogen

volume of 30 moles of hydrogen = 30 × 24 = 720 dm^3

Notice that there is a much quicker and easier way of doing this calculation. For gases only, the volume is directly proportional to the number of moles.
This means that:

	$N_2(g)$	+	$3H_2(g)$	→	$2NH_3(g)$
mole ratio	1	:	3	:	2
volume ratio	1	:	3	:	2
volume in dm^3	240	:	720	:	480

Percentage yield

If the reactants shown in an equation are converted completely into the products, we say that the **percentage yield** is 100%. However, in some circumstances, yields are less than 100%.

Skills

Calculating percentage yield

You can calculate percentage yield using the equation:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Worked example

0.60 g of magnesium ribbon, Mg, was burned in excess oxygen, O_2, according to the equation:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

The mass of magnesium oxide, MgO, that was produced was found to be 0.80 g.

Calculate the percentage yield.

A_r: Mg = 24, O = 16

Answer

M_r of MgO = 24 + 16 = 40

moles of Mg = 0.60 + 24 = 0.025

The mole ratio from the equation is 1 mole of Mg produces 1 mole of MgO.

Therefore, 0.025 moles of Mg produce 0.025 moles of MgO.

mass of MgO = 0.025 × 40 = 1.00 g

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This is the theoretical yield: 1.00 g of MgO would be produced if the percentage yield was 100%.

However, the actual yield is only 0.80 g, so:

$$\text{percentage yield} = \frac{0.80}{1.00} \times 100 = 80.0\%$$

Percentage purity

Naturally occurring substances are impure – a sample will contain other substances mixed with the one that we are interested in. The percentage (by mass) of the element or compound we want in a sample is known as the **percentage purity.**

Skills

Calculating percentage purity

We can calculate percentage purity using the following equation:

$$\text{percentage purity} = \frac{\text{mass of the pure product}}{\text{mass of the impure product obtained}} \times 100$$

This means we first need to find out the mass of substance in a sample. We can use ideas about moles and chemical equations to analyse the result of an experiment.

Worked example

Limestone is often thought of as calcium carbonate, $CaCO_3(s)$. However, most limestones contain other compounds, too.

1.00 g of limestone from a quarry is added to 100 cm^3 of 0.200 mol/dm^3 hydrochloric acid, $HCl(aq)$ (in excess):

Equation 1: $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

The leftover acid is titrated and found to be neutralised by 24.8 cm^3 of 0.100 mol/dm^3 sodium hydroxide solution, $NaOH(aq)$:

Equation 2: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Calculate the percentage purity of the limestone from the quarry.

Answer

Moles of NaOH that reacted with the leftover acid:

$$\text{number of moles} = \frac{\text{volume (in cm}^3) \times \text{concentration (in mol/dm}^3)}{1000}$$

$$= \frac{2.48\ \text{cm}^3 \times 0.100\ \text{mol/dm}^3}{1000}$$

$$= 2.48 \times 10^{-3}\ \text{moles}$$

The mole ratio in equation 2 is 1 mole of NaOH reacts with 1 mole of HCl.

So, 2.48×10^{-3} moles of NaOH react with 2.48×10^{-3} moles of HCl.

So, 2.48×10^{-3} moles of HCl are left over.

Moles of HCl that were added to the limestone

$$= \frac{100\ \text{cm}^3 \times 0.2\ \text{mol/dm}^3}{1000} = 0.02\ \text{moles}$$

Moles of HCl that reacted with calcium carbonate, $CaCO_3$

= moles of HCl added to limestone − moles HCl left over

$= 0.02 - 2.48 \times 10^{-3} = 0.01752$ moles HCl

The mole ratio in equation 1 is 2 moles of HCl react with 1 mole of $CaCO_3$.

So, 0.01752 moles of HCl react with 0.01752 ÷ 2 = 0.00876 moles of $CaCO_3$.

M_r of $CaCO_3 = 40 + 12 + (16 \times 3) = 100$

mass of $CaCO_3$ = number of moles × molar mass

$= 0.00876 \times 100 = 0.876$ g

percentage of $CaCO_3$ in limestone

$$= \frac{\text{mass of CaCO}_3}{\text{mass of limestone}} \times 100$$

$$= \frac{0.876}{1.00} \times 100 = 87.6\%$$

So, the percentage purity is 87.6%.

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Percentage composition

Iron(II) sulfide has the formula FeS.

The percentages of iron and sulfur in iron(II) sulfide are:

$$Fe = (56 \div 88) \times 100 = 63.6\%$$

$$S = (32 \div 88) \times 100 = 36.4\%$$

This means that all samples of iron(II) sulfide contain 63.6% iron and 36.4% sulfur by mass.

Limiting reactants

When two substances are mixed, you cannot assume that both substances will react completely and that neither is left over. This is possible, but it is also possible that too much of one substance is used and some of it will be left over at the end of the reaction. The substance that is all used up is called the **limiting reactant** and the other substance is said to be **in excess**.

Skills

Which is the limiting reactant?

We can use equations and ideas about moles to work out which substance in a reaction limits the amount of product formed.

Worked example

5.6 g of iron, Fe, and 4.0 g of sulfur, S, are mixed together and heated. The equation for the reaction is:

$$Fe(s) + S(s) \rightarrow FeS(s)$$

Deduce which substance is the limiting reactant.

Answer

moles of Fe = 5.6 ÷ 56 = 0.10

moles of S = 4.0 ÷ 32 = 0.125

The mole ratio from the equation is 1 mole of Fe reacts with 1 mole of S.

Therefore, 0.10 moles of Fe react with 0.10 moles of S.

However, there are 0.125 moles of S.

0.125 is greater than 0.10, therefore some S is left over.

Therefore, S is in excess and Fe is the limiting reactant.

Revision activity

Use the worked examples to make a flash card for each type of calculation in this chapter, so you can test yourself on how to do them. On the front of each card, write the type of information you are given and what you need to calculate. On the back of each card, write the steps you need to take to complete the calculation. You may like to add icons or use colour coding for the different types of information or/and different steps.

Sample questions

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1 Calculate the concentration in g/dm^3 of a solution containing:

a 20.0 g of NaOH in 500 cm^3 [1]

b 17.4 g of K_2SO_4 in 2 dm^3 [1]

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Student's answers

a 20.0 ÷ 500 = 0.04 g/dm^3
b 17.4 ÷ 174 = 0.1
0.1 ÷ 2 = 0.05 g/dm^3

Teacher's comments

Units are extremely important in all calculations.

This question asks for both answers in g/dm^3.

a The student has not converted cm^3 into dm^3. The answer is therefore in g/cm^3.
b The student has converted grams into moles by dividing by 174, which is the M_r of K_2SO_4. The answer is therefore in mol/dm^3.

Correct answers

a (20.0 × 1000) ÷ 500 = 40.0 g/dm^3
b 17.4 ÷ 2 = 8.7 g/dm^3

2 A compound has the following percentage composition by mass:
43.7 % phosphorus 56.3% oxygen
Calculate the empirical formula of the compound.
A_r: P = 31, O = 16

Student's answer

moles of P = 43.7 ÷ 31 = 1.41

moles of O = 56.3 ÷ 16 = 3.52

Dividing both by the smallest:

1.41 ÷ 1.41 = 1

3.52 ÷ 1.41 = 2.5 ≈ 3

Therefore, the empirical formula is PO_3.

Correct answer

100 g of the compound contains 43.7 g of phosphorus, P, and 56.3 g of oxygen, O.

moles of P = 43.7 ÷ 31 = 1.41

moles of O = 56.3 ÷ 16 = 3.52

Dividing both by the smallest:

1.41 ÷ 1.41 = 1

3.52 ÷ 1.41 = 2.5

Multiplying both by 2:

1 × 2 = 2

2.5 × 2 = 5

The empirical formula is P_2O_5.

Teacher's comments

It is a common error to say 2.5 ≈ 3 or, alternatively, 2.5 ≈ 2. Both are incorrect.

If dividing the number of moles of atoms by the smallest does not produce whole numbers in all cases, the answers should be multiplied by 2. If the answers are still not whole numbers, multiply by increasing whole numbers until whole numbers are achieved in all cases.

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3 Calcium carbonate decomposes when it is heated according to the equation:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Calculate the mass of calcium oxide, CaO, that is produced when 20.0 g of calcium carbonate, $CaCO_3$, is heated until there is no further change.

Student's answer

M_r of $CaCO_3$ = 20 + 6 + (8 × 3) = 50

M_r of CaO = 20 + 8 = 28

M_r of CO_2 = 6 + (8 × 2) = 22

moles of $CaCO_3$ = 20.0 ÷ 50 = 0.40

moles of CaO = 0.40

mass of CaO = 0.40 × 28 = 11.2

Correct answer

M_r of $CaCO_3$ = 40 + 12 + (16 × 3) = 100

M_r of CaO = 40 + 16 = 56

moles of $CaCO_3$ = 20.0 ÷ 100 = 0.20 moles

The mole ratio from the equation is 1 mole of $CaCO_3$: 1 mole of CaO.

0.20 moles $CaCO_3$: 0.20 moles of CaO

mass of CaO = number of moles × molar mass

= 0.20 × 56 = 11.2 g

Teacher's comments

The student used proton numbers instead of A_r values to calculate the M_r values. By luck, this meant that the final answer of 11.2 was correct. However, the student would *not* have achieved full marks.

The question does not ask about carbon dioxide, CO_2, so there is no need to calculate the relative molecular mass, M_r, of carbon dioxide.

The student should have given the answer in the correct units, g.

Exam-style questions

1 Calculate the M_r of the following compounds:

- **a** glucose, $C_6H_{12}O_6$ [1]
- **b** hydrated sodium sulfate, $Na_2SO_4.10H_2O$ [1]
- **c** potassium dichromate(VI), $K_2Cr_2O_7$ [1]

[Total: 3]

2 6.0 g of magnesium ribbon burns in excess oxygen to form 10.0 g of magnesium oxide.

- **a** State what is meant by *excess oxygen* in the statement above. [1]
- **b** State the mass of magnesium oxide that would form if 18.0 g of magnesium ribbon was burnt in excess oxygen. [1]
- **c** State the mass of magnesium ribbon that would be burned in excess oxygen to form 0.24 g of magnesium oxide. [1]

[Total: 3]

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In the questions which follow, use the following values of A_r:

H = 1; C = 12; O = 16; Na = 23, A*l* = 27; S = 32, C*l* = 35.5; K = 39; Ca = 40; Ti = 48

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure.

3 Calculate the mass of hydrogen gas produced when 8.1 g of aluminium powder reacts with excess dilute hydrochloric acid, HC*l*, according to the equation:

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$ [Total: 3]

4 Calculate the mass of calcium carbide, $CaC_2(s)$, required to produce 120 cm^3 of ethyne gas, $C_2H_2(g)$, by reaction with excess water according to the equation:

$$CaC_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$$ [Total: 3]

5 20.0 cm^3 of aqueous potassium hydroxide, KOH(aq), neutralised 35.0 cm^3 of dilute sulfuric acid, $H_2SO_4(aq)$, whose concentration was 0.20 mol/dm^3. The equation is:

$$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

Calculate the concentration of the aqueous potassium hydroxide, KOH(aq), in:

a mol/dm^3

b g/dm^3 [Total: 4]

6 A compound has composition by mass which is 54.5% carbon, 9.1% hydrogen and 36.4% oxygen.
The M_r of the compound is 44. Calculate the:

a empirical formula

b molecular formula of the compound. [Total: 4]

7 When 0.38 g of titanium(IV) chloride, $TiCl_4(l)$, reacted with excess sodium, the reaction produced 0.024 g of titanium, Ti(s). The equation is:

$$TiCl_4(l) + 4Na(s) \rightarrow Ti(s) + 4NaCl(s)$$

Calculate the percentage yield of titanium. [Total: 4]

8 2.3 g of sodium, Na(s), is burned in 7.1 g of chlorine, Cl_2.
Sodium chloride, NaC*l*(s), is the only product. The equation is:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

a Deduce which reactant is the limiting reactant. Explain your answer. [2]

b Calculate the mass of sodium chloride which forms. [3]

[Total: 5]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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5 Electrochemistry

Key objectives

By the end of this section, you should be able to:

- define electrolysis
- in a simple electrolytic cell, identify:
  - the anode
  - the cathode
  - the electrolyte
- describe the transfer of charge during electrolysis to include:
  - the movement of electrons in the external circuit
  - the loss or gain of electrons at the electrodes
  - the movement of ions in the electrolyte
- describe the products of electrolysis and state the observations made during the electrolysis of:
  - molten lead(II) bromide
  - concentrated aqueous sodium chloride
  - dilute sulfuric acid
- state that metals or hydrogen are formed at the cathode and that non-metals (other than hydrogen) are formed at the anode
- predict the identity of the products at each electrode for the electrolysis of a molten binary compound
- identify the products formed at the electrodes and describe the observations made during the electrolysis of:
  - aqueous copper(II) sulfate using inert carbon/graphite electrodes
  - aqueous copper(II) sulfate using copper electrodes
- predict the identity of the products formed at each electrode for the electrolysis of a dilute or concentrated aqueous solution of a halide
- construct ionic half-equations for reactions at the anode (to show oxidation) and at the cathode (to show reduction)
- state that the main ore of aluminium is bauxite and that aluminium is extracted by electrolysis
- describe the extraction of aluminium from bauxite, including the role of cryolite and the reactions at the electrodes
- state that a hydrogen–oxygen fuel cell uses hydrogen and oxygen to produce electricity, with water as the only chemical product
- describe the advantages and disadvantages of using hydrogen–oxygen fuel cells in comparison with gasoline/petrol engines in vehicles
- know why metal objects are electroplated
- describe how metals are electroplated

Key terms

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Term	Definition
Anion	A negative ion. Anions are attracted to the anode in electrolysis.
Anode	The positive (+) electrode. It is positively charged because electrons are drawn away from it.
Binary compound	A compound containing two elements chemically combined.
Cathode	The negative (–) electrode. It is negatively charged because an excess of electrons move towards it.
Cation	A positive ion. Cations are attracted to the cathode in electrolysis.
Electrodes	The conducting rods by which electric current enters and leaves the electrolyte.
Electrolysis	The decomposition of an ionic compound, when molten or in aqueous solution, by the passage of an electric current.
Electrolyte	A liquid which will carry electric current and is chemically changed by it.
Inert electrode	An electrode that does not react with the electrolyte or the products of electrolysis. Examples are carbon (graphite) and platinum.

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5.1 Electricity and chemistry

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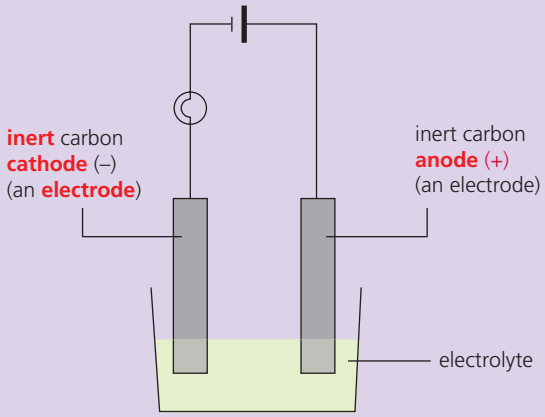
Electrolysis is the decomposition of a molten ionic compound, or an ionic compound dissolved in water, by the passage of an electric current.

Skills

Laboratory electrolysis

The liquid in electrolysis is known as the **electrolyte**. It is placed in a crucible (if it is a solid that needs to be heated to its melting point) or a beaker (if it is a liquid at room temperature).

Figure 5.1 shows a commonly used arrangement and some key terms.



▲ **Figure 5.1 The important terms used in electrolysis**

Electrolytes

Substances that conduct electricity can be subdivided into conductors and electrolytes.

▼ **Table 5.1 Differences between conductors and electrolytes**

	Conductors	Electrolytes
Physical state	Solid	Liquid
Differences	They conduct electricity but are *not* chemically changed by the electric current. They become hot, which is a physical change.	They conduct electricity and are chemically changed by the electric current. The products of the chemical change are formed at the electrodes.
Examples	All metallic elements and alloys Graphite and graphene	Molten ionic compounds Aqueous solutions containing ions
Particles responsible for conduction	Moving (mobile) electrons	Moving ions

Electrolytes must be in the liquid state.

Solid ionic compounds, such as sodium chloride, do not conduct electricity because the oppositely charged ions are held together in the giant ionic lattice by strong electrostatic attraction. As the ions are not moving, solid sodium chloride does not conduct electricity.

Two ways to make ionic solids into electrolytes

1. Melt the solid. This requires a large amount of heat energy because ionic compounds have high melting points (see Chapter 3). Molten ionic compounds are electrolytes because ions are moving when the compound is in the liquid state.
2. Dissolve the solid in water. An aqueous solution of an ionic compound also contains moving ions.

Changes at the electrodes

When electrolytes conduct electricity, the positive ions (**cations**) move to the **cathode** (–) and the negative ions (**anions**) move to the **anode** (+).

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The products of electrolysis are formed at the electrodes.

When ions *lose* their charge to form atoms or molecules, they are said to be *discharged*.

At the cathode, positive ions *gain* electrons and are *reduced*. For example:

$$2H^+ + 2e^- \rightarrow H_2$$

H^+ ions are reduced because:

- they gain electrons
- there is a decrease in the oxidation number from +1 to 0

At the anode, negative ions *lose* electrons and are *oxidised*. For example:

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

Cl^- ions are oxidised because:

- they lose electrons
- there is an increase in the oxidation number from –1 to 0

5.2 Electrolysis of molten lead(II) bromide

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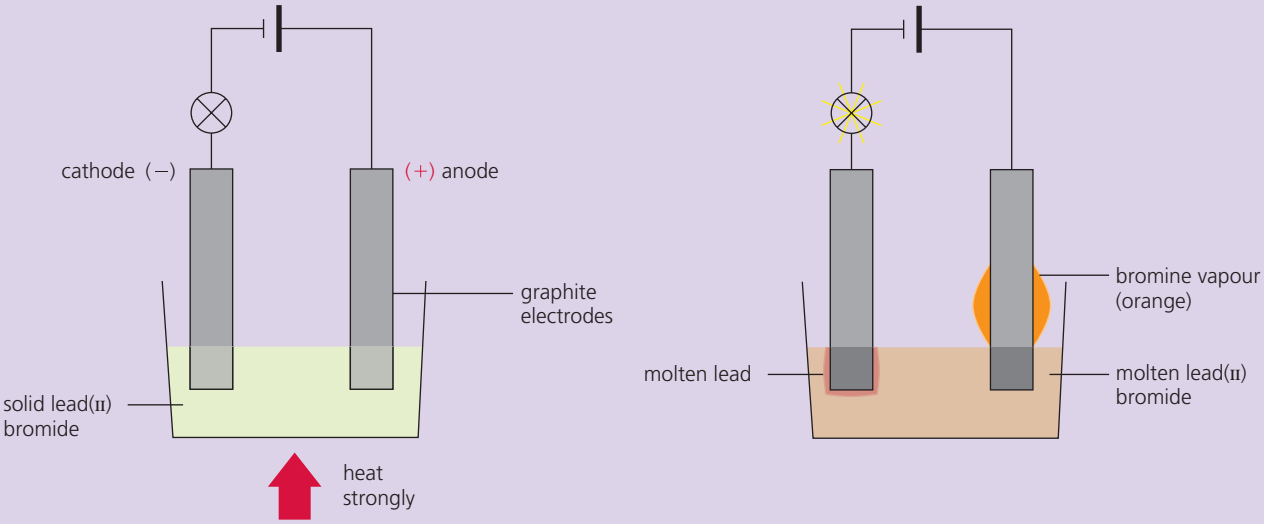
Skills

Electrolysis of molten lead(II) bromide

Lead(II) bromide is an ionic solid. It is not an electrical conductor in the solid state. It is insoluble in water. Thus, it only undergoes electrolysis when molten.



▲ **Figure 5.2 Electrolysis of molten lead(II) bromide**

At the high temperature of the electrolysis, both bromine and lead will vaporise. Therefore, the lead(II) bromide must be electrolysed in a fume cupboard.

The products of this electrolysis are:

- bromine, which is seen at the anode as an orange–red vapour
- lead, which is seen at the cathode (after cooling) as a silvery grey metal

Electrolysis of molten binary compounds

When molten **binary** ionic compounds are electrolysed, the non-metallic element is formed at the positive electrode (anode) and the metallic element is formed at the negative electrode (cathode). Some examples are shown in Table 5.2.

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▼ **Table 5.2 Products of electrolysis of molten binary compounds**

Electrolyte	Product at anode (+)	Product at cathode (–)
Molten lead bromide	Bromine	Lead
Molten potassium iodide	Iodine	Potassium
Molten sodium chloride	Chlorine	Sodium

Ionic half-equations at the electrodes

At the anode, negative ions are oxidised by losing electrons.

At the cathode, positive ions are reduced by gaining electrons.

The ionic half-equations for the examples in Table 5.2 are shown in Table 5.3.

▼ **Table 5.3 Ionic half-equations for electrolysis of molten binary compounds**

Electrolyte	Reaction at anode (+)	Reaction at cathode (–)
Molten lead bromide	$2Br^- \rightarrow Br_2 + 2e^-$	$Pb^{2+} + 2e^- \rightarrow Pb$
Molten potassium iodide	$2I^- \rightarrow I_2 + 2e^-$	$K^+ + e^- \rightarrow K$
Molten sodium chloride	$2Cl^- \rightarrow Cl_2 + 2e^-$	$Na^+ + e^- \rightarrow Na$

Revision activity

Draw a series of pictures, like a comic strip, to show how ions move in an electrolyte and what happens to the ions at each electrode. If you create your pictures using software, or if you photograph or scan your pictures, you may like to create a gif to share.

5.3 Electrolysis of aluminium oxide

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Aluminium is extracted from bauxite, which is impure aluminium oxide, Al_2O_3. Bauxite is first purified and then electrolysis is carried out.

Electrolysis has to be used because aluminium oxide is not reduced by carbon monoxide or any other common reducing agent.

This process is expensive due to the high cost of electricity.

Aluminium oxide has a melting point of 2071°C. To achieve such a high temperature would require a large amount of heat energy and increase costs further.

So the aluminium oxide is first dissolved in another aluminium compound, molten cryolite, Na_3AlF_6. The advantages of this are that:

- the electrolyte can be maintained in the liquid state between 800°C and 1000°C, a temperature considerably lower than 2071°C, which greatly reduces energy costs
- cryolite improves the conductivity of the electrolyte.

Aluminium oxide in molten cryolite behaves in the same way as molten aluminium oxide as far as the products of electrolysis are concerned.

Electrolysis is carried out in a steel tank using carbon (graphite) as electrodes. The anodes are carbon (graphite) blocks which are lowered into the electrolyte. The cathode is the carbon (graphite) lining of the tank.

The electrode reactions are:

cathode (–) $Al^{3+} + 3e^- \rightarrow Al$

anode (+) $2O^{2-} \rightarrow O_2 + 4e^-$

Molten aluminium collects at the bottom of the tank and is siphoned off.

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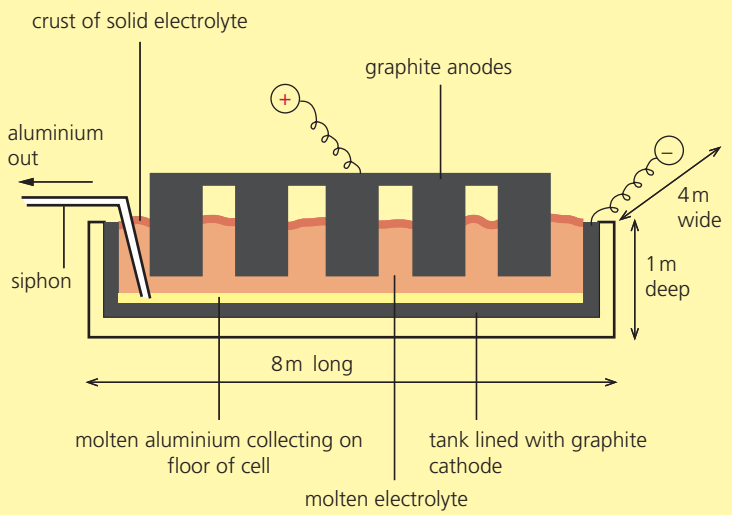

▲ **Figure 5.3 The Hall–Héroult cell is used in industry to extract aluminium**

The oxygen that is produced at the anode reacts, at the high temperature of the cell, with the graphite anodes, producing carbon dioxide gas which escapes:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Thus, the anodes burn away and need to be replaced regularly.

The cost of electricity is the largest expense in this process, so it is carried out in regions where cheap electricity is available, for example from hydroelectric power.

5.4 Electrolysis of aqueous solutions

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Skills

Electrolysis of dilute sulfuric acid

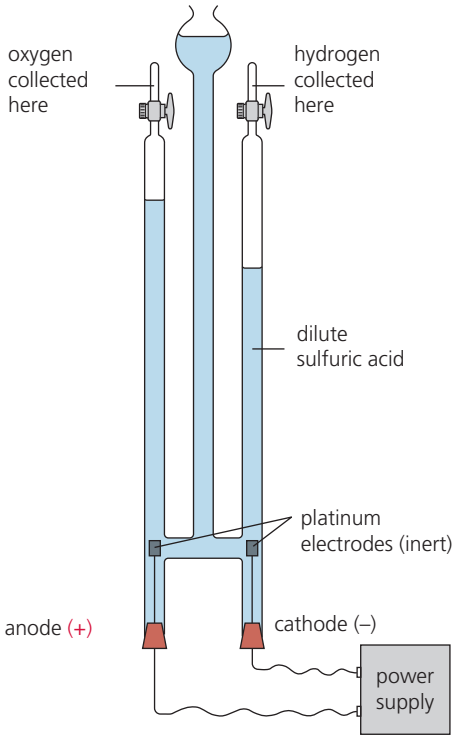
The gases produced in this process can be collected using a Hofmann voltameter (Figure 5.4) or the alternative apparatus shown in Figure 5.5 on page 54.

The gases collected can then be tested (see Chapter 14).

Products of electrolysis

Molten ionic compounds produce a non-metallic element at the anode and a metallic element at the cathode.

Aqueous solutions produce oxygen or a halogen at the anode and hydrogen or a metal at the cathode. The hydrogen and oxygen come from the water that is contained in the aqueous solution.



▲ **Figure 5.4 A Hofmann voltameter used to electrolyse dilute sulfuric acid**

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▼ **Table 5.4 Summary of products formed during electrolysis.**

Type of electrolyte	Products at anode (+)	Products at cathode (–)
Molten ionic compound	Non-metallic element	Metallic element
Aqueous solution containing ions	Either oxygen gas or if the electrolyte is a concentrated solution of a halide (chloride, bromide or iodide), a halogen (chlorine, bromine or iodine)	Either hydrogen gas or metallic element below hydrogen in the reactivity series (e.g. copper)

The products, observations and half-equations for the electrolysis of different electrolytes when using inert carbon or platinum electrodes are shown in Table 5.5.

▼ **Table 5.5 Examples of products of the electrolysis of different electrolytes, using inert electrodes**

	Product at anode (+)	Observations at anode (+)	Reaction at anode (+)	Product at cathode (–)	Observations at cathode (–)	Reaction at cathode (–)
Molten sodium chloride, $NaCl(l)$	Chlorine	Bubbles of green gas	$2Cl^- \rightarrow Cl_2 + 2e^-$	Sodium	Grey metal coating	$Na^+ + e^- \rightarrow Na$
Concentrated aqueous sodium chloride, $NaCl(aq)$	Chlorine	Bubbles of green gas	$2Cl^- \rightarrow Cl_2 + 2e^-$	Hydrogen	Bubbles of colourless gas	$2H^+ + 2e^- \rightarrow H_2$
Molten lead bromide, $PbBr_2(l)$	Bromine	Bubbles of brown gas	$2Br^- \rightarrow Br_2 + 2e^-$	Lead	Grey metal coating	$Pb^{2+} + 2e^- \rightarrow Pb$
Concentrated hydrochloric acid, $HCl(aq)$	Chlorine	Bubbles of green gas	$2Cl^- \rightarrow Cl_2 + 2e^-$	Hydrogen	Bubbles of colourless gas	$2H^+ + 2e^- \rightarrow H_2$
Dilute sulfuric acid, $H_2SO_4(aq)$	Oxygen	Bubbles of colourless gas	$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$	Hydrogen	Bubbles of colourless gas	$2H^+ + 2e^- \rightarrow H_2$
Aqueous copper(II) sulfate, $CuSO_4(aq)$	Oxygen	Bubbles of colourless gas	$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$	Copper	Pink metal coating	$Cu^{2+} + 2e^- \rightarrow Cu$

Make sure you remember:

- Aqueous solutions of acids always produce hydrogen at the cathode. The H^+ ion is found in both the acidic substance and the water.
- During the electrolysis of any aqueous solution containing positive ions of a metal *above hydrogen* in the reactivity series, hydrogen is produced at the cathode, *not* the metallic element.
- Very reactive metals that react with cold water (such as potassium, sodium and calcium) *cannot* be produced by electrolysis of aqueous solutions containing ions of these metals. These metals can only be extracted by electrolysis using a molten electrolyte (see Section 5.3).

Revision activity

There are many similar words in this topic that are easily confused. Draw a concept map by writing the key words on a large sheet of paper. Space them out well. Then add labelled arrows to show how the words are linked to each other.

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5.5 Electrolysis of copper(II) sulfate aqueous solution

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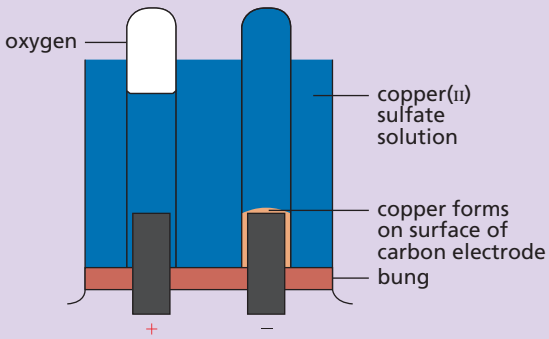
Skills

Electrolysis of copper(II) sulfate aqueous solution

If aqueous copper(II) sulfate is electrolysed using carbon or platinum electrodes (inert electrodes), the products are copper at the cathode and oxygen at the anode (see Table 5.5).



▲ **Figure 5.5 The electrolysis of copper(II) sulfate solution using inert electrodes**

The solution gradually loses its blue colour and eventually turns colourless. The blue colour is due to $Cu^{2+}(aq)$ and the colour gradually fades because the concentration of $Cu^{2+}(aq)$ decreases.

However, if the anode is made of copper, oxygen is not produced at the anode. Instead, the copper anode goes into solution as positive ions:

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$$

The mass of the anode will decrease, but this change may not be visible over a short period of time.

Eventually, the $Cu^{2+}(aq)$ ions reach the cathode, where the reverse reaction occurs and copper metal is formed:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

The mass of the cathode will gradually increase as fresh copper is deposited.

As each $Cu^{2+}(aq)$ ion that is removed from the solution at the cathode is replaced by a $Cu^{2+}(aq)$ ion forming at the anode, the blue colour of the solution does not change.

5.6 Fuel cells

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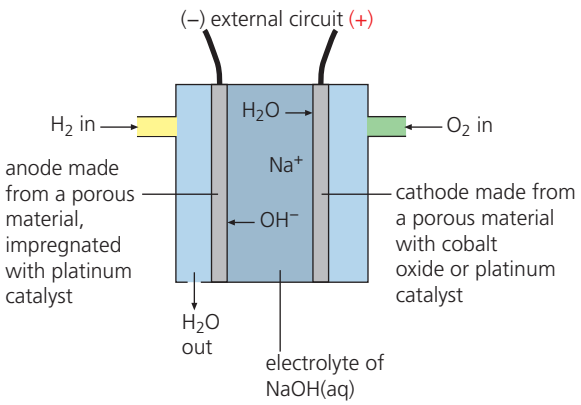
Hydrogen–oxygen fuel cells are used in electric cars. Hydrogen is used as a fuel as an alternative to petrol (gasoline).

The overall reaction is the same as when hydrogen is burned in air or oxygen:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

However, the hydrogen does not undergo combustion – it reacts with oxygen to produce electricity.

Fuel cells operate in acidic or alkaline conditions. An alkaline hydrogen fuel cell is shown in Figure 5.6.



▲ **Figure 5.6 A diagrammatic view of a fuel cell**

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Advantages and disadvantages

Advantages of using hydrogen–oxygen fuel cells rather than petrol engines are:

- Fuel cells are much more efficient than internal combustion engines, which means there is much less energy loss.
- Fuel cells produce water as the only chemical product. Petrol engines produce carbon dioxide, which is a greenhouse gas (see Chapter 11).

Disadvantages of using fuel cells rather than petrol engines are:

- Hydrogen is a gas and is difficult to store.
- Gaseous hydrogen has a low density. Therefore, a tank of hydrogen contains a much lower mass than one of the same size filled with petrol. Thus, less energy will be produced.
- The infrastructure to enable widespread use of electric cars will cost a lot to develop.

Revision activity

Draw a table to show the advantages and disadvantages of using hydrogen–oxygen fuel cells in cars rather than a petrol or diesel engine.

5.7 Electroplating

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Electroplating is an application of electrolysis that can be carried out in a school laboratory or on a large scale. Electroplating means coating a metal with a thin layer of another metal.

The purposes are:

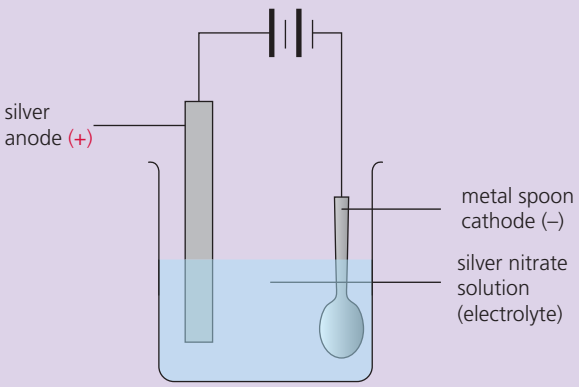
- to improve appearance
- to prevent corrosion, for example to prevent rusting of iron or steel

Skills

Electroplating

Electroplating is carried out using:

- the plating metal as the anode
- the object to be plated as the cathode
- an aqueous solution containing ions of the plating metal as the electrolyte



▲ **Figure 5.7 Silver plating a spoon**

In the example shown in Figure 5.7:

- the plating metal (anode) is silver
- the object to be plated (cathode) is a metal spoon
- the electrolyte is an aqueous solution of silver nitrate containing $Ag^+(aq)$

The silver produced at the cathode electroplates the spoon.

The silver ions that are released at the anode replace those that are discharged.

The silver ions in the electrolyte are discharged at the cathode:

$$Ag^+(aq) + e^- \rightarrow Ag(s)$$

The silver anode goes into solution as silver ions:

$$Ag(s) \rightarrow Ag^+(aq) + e^-$$

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Sample questions

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1 State what is meant by the term *electrolysis*. [2]

Student's answer

Electrolysis is the breaking of a compound using electricity.

Teacher's comments

There are four important parts to the correct answer:
- decomposition
- ionic compound
- molten or aqueous solution
- electricity

The student:
- uses the term *breaking*, which is not the same as decomposition: breaking is considered to refer to a physical change, as in breaking something into smaller pieces
- does not mention that the compound is ionic
- does not refer to molten or aqueous solution
- correctly refers to electricity

Correct answer

Electrolysis is the decomposition of an ionic compound, when molten or in aqueous solution, by the passage of an electric current.

2 Complete the table below.

Electrolyte	Product at anode (+)	Product at cathode (–)
Molten lead(II) bromide		
Concentrated aqueous sodium chloride		
Dilute sulfuric acid		

Student's answers

Electrolyte	Product at anode (+)	Product at cathode (–)
Molten lead(II) bromide	Bromine	Lead
Concentrated aqueous sodium chloride	Oxygen	Sodium
Dilute sulfuric acid	Hydrogen	Oxygen

Teacher's comments

The products of electrolysis of molten lead(II) bromide are correct and are placed at the correct electrodes.

The electrolysis of *concentrated* aqueous sodium chloride produces chlorine at the anode. Oxygen would only be produced at the anode if the solution was *dilute*. It is a very common error to state that sodium is produced at the cathode. Reactive metals, such as sodium, are only produced at the

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cathode using a molten sodium compound as the electrolyte. If an aqueous solution is used, hydrogen is produced at the cathode.

The products of electrolysis of dilute sulfuric acid are correct but they are placed at the incorrect electrodes. Getting the two electrodes the wrong way round like this leads to the loss of many of the marks available. It is therefore extremely important to learn that the anode is positive (+) and the cathode is negative (–).

Correct answers

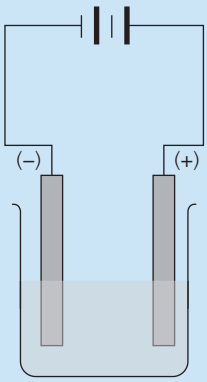
Electrolyte	Product at anode (+)	Product at cathode (–)
Molten lead(II) bromide	Bromine	Lead
Concentrated aqueous sodium chloride	Chlorine	Hydrogen
Dilute sulfuric acid	Oxygen	Hydrogen

Exam-style questions

1 Complete the following table to show the products of electrolysis using carbon/graphite electrodes. [Total: 7]

Electrolyte	Name of product at anode (+)	Name of product at cathode (–)
Molten potassium bromide	[1]	[1]
Molten sodium chloride	[1]	[1]
Concentrated aqueous sodium chloride	[1]	[1]
[1]	Iodine	Lead

2 Aqueous lithium chloride is electrolysed using the apparatus opposite.



a On a copy of the diagram, label:
- **i** the anode [1]
- **ii** the cathode [1]
- **iii** the electrolyte [1]

b State what is meant by the term *aqueous*. [1]

c Explain why aqueous lithium chloride is able to conduct electricity. [2]

[Total: 6]

3 Aluminium is extracted by electrolysis from its main ore.

a Name the main ore of aluminium from which aluminium is extracted. [1]

b Use your knowledge of the reactivity series to explain why aluminium is not extracted by reduction of its oxide using carbon. [1]

[Total: 2]

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4 Use the following words to fill in the spaces in the passage that follows. Each word should be used once.

anode **appearance** **cathode** **corrosion**
electrodes **electrolysis** **electrolyte**

_____ is the process in which an _____ is decomposed. The products of decomposition are formed at the _____ . The positive electrode is called the _____ and the negative electrode is called the _____ .
Electroplating means covering a metal object with a thin layer of another metal. One of the reasons for electroplating is to improve _____ . Another reason is to resist _____ . [Total: 7]

5 A student wanted to electroplate a knife with nickel. What should the student use as:

a the anode [1]
b the electrolyte [1]
c the cathode? [1]

[Total: 3]

6 A student carries out electrolysis of concentrated aqueous potassium iodide in a beaker using carbon electrodes.

a Name the product at the anode. [1]
b Write an ionic half-equation for the reaction occurring at the cathode. [1]
c State the type of reaction occurring at the anode. [1]
d State the name of the solution left in the beaker when the electrolysis has finished. [1]
e Name the type of particles that are responsible for the conduction of electricity in the conducting wire. [1]
f Name the type of particles that are responsible for the conduction of electricity in the electrolyte. [1]

[Total: 6]

7 **a** Complete the table below to show the products of electrolysis using carbon/graphite electrodes.

Electrolyte	Name of product at anode (+)	Name of product at cathode (–)
Aqueous copper(II) sulfate	[1]	[1]
Concentrated aqueous lithium bromide	[1]	[1]
Dilute aqueous sodium chloride	[1]	[1]

b Describe how the colour of aqueous copper(II) sulfate solution changes when it undergoes electrolysis using carbon/graphite electrodes. [2]

c If aqueous copper(II) sulfate undergoes electrolysis using copper electrodes, state what the difference would be in:

i the change that occurs at the anode [1]
ii the observation made in the aqueous solution [1]

[Total: 10]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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6 Chemical energetics

Key objectives

By the end of this section, you should be able to:

- name the fossil fuels: coal, natural gas and petroleum
- name methane as the main constituent of natural gas
- state that petroleum is a mixture of hydrocarbons
- describe the separation of petroleum into useful fractions
- describe how the properties of fractions of petroleum change from the bottom to the top of the fractionating column
- name the uses of the fractions
- state that an exothermic reaction transfers heat energy to the surroundings leading to an increase in the temperature of the surroundings
- state that an endothermic reaction takes in heat energy from the surroundings leading to a decrease in the temperature of the surroundings
- interpret reaction pathway diagrams showing exothermic and endothermic reactions

- state that the transfer of thermal energy during a chemical reaction is called the enthalpy change, ΔH, of the reaction
- define activation energy, E_a
- draw and label reaction pathway diagrams for exothermic and endothermic reactions using information provided, to include:
  - reactants
  - products
  - enthalpy change, ΔH
  - activation energy, E_a
- state that bond breaking is an endothermic process and that bond making is an exothermic process
- explain the enthalpy change of a reaction in terms of bond breaking and bond making
- calculate the enthalpy change of a reaction using bond energies

Key terms

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Term	Definition
Activation energy	The activation energy, E_a, is the minimum energy that colliding particles must have in order to react.
Bond energy	Amount of energy required to break one mole of covalent bonds in gaseous molecules.
Endothermic reaction	An endothermic reaction absorbs thermal energy from the surroundings leading to a decrease in temperature of the surroundings.
Enthalpy change	The transfer of thermal energy during a reaction is called the enthalpy change, ΔH, for the reaction. ΔH is negative for exothermic reactions and positive for endothermic reactions.
Exothermic reaction	An exothermic reaction transfers thermal energy to the surroundings leading to an increase in temperature of the surroundings.
Fossil fuels	Fuels, such as coal, petroleum and natural gas, formed from the remains of plants and animals.
Fractional distillation	A technique used to separate a mixture of liquids that have different boiling points.
Fuel	A substance that can be conveniently used as a source of energy.
Oil refining	The process of converting petroleum into separate fractions.

6.1 Substances from petroleum

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Petroleum (crude oil) is a mixture of hydrocarbons (see Chapter 12). Separating it by **fractional distillation** gives mixtures of hydrocarbons with a narrow range of boiling points. These mixtures are called **fractions**.

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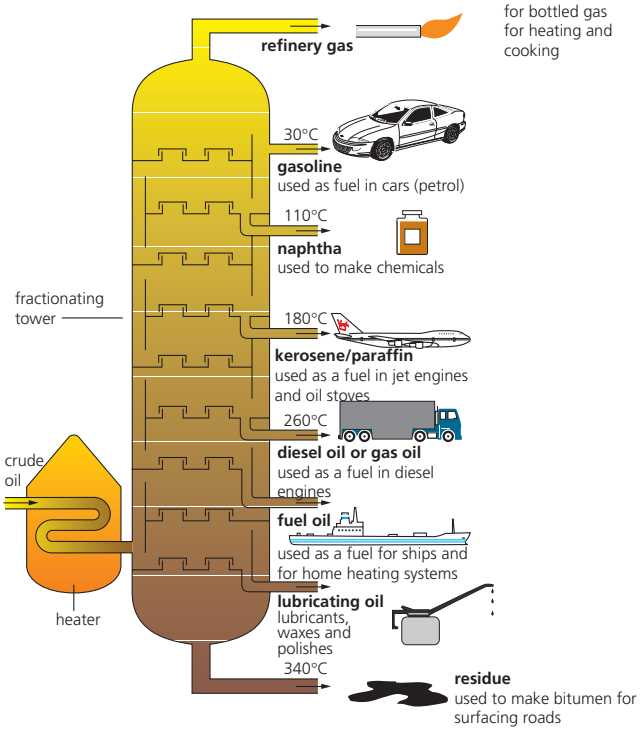
Properties of the fractions

From the bottom to the top of the fractionating column, the properties of the fractions change in the following ways:

- The chain length decreases.
- The volatility (how easily they evaporate) gets higher.
- The boiling point gets lower.
- The viscosity (stickiness) gets lower.

Uses of the fractions

The uses of the fractions are shown in Figure 6.1.



▲ **Figure 6.1 Uses of the different fractions obtained from crude oil**

Revision activity

Use the information above to create a table which links the properties of fractions and the way the properties change from the bottom to the top of the fractionating column.

Revision activity

Make two sets of 8 cards each.

- Set 1: Names of fractions, e.g. fuel oil, gasoline etc.
- Set 2: Uses of fractions, e.g. surfacing roads, heating etc.

Shuffle the cards in each set. Then try to match the name of each fraction with its use. You could do this with a friend.

6.2 What is a fuel?

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A **fuel** is any substance which can be used as a source of energy. Fossil fuels release energy in the form of heat when they undergo combustion.

6.3 Fossil fuels

REVISED

Fossil fuels are fuels formed by natural processes over millions of years as a result of the decay of buried dead organisms. Examples are coal, natural gas and petroleum (crude oil). Fossil fuels are a finite resource because once they run out, they cannot be replaced. They are *non-renewable*.

Methane, CH_4, is the main constituent of natural gas.

6.4 Alternatives to fossil fuels

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Alternatives to fossil fuels are sources of energy, such as:

- nuclear fuels
- biomass and biogas
- hydrogen
- hydroelectric power
- wind
- solar energy

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6.5 Exothermic and endothermic reactions

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Energy changes in reactions

Exothermic reactions are reactions in which thermal energy is given out to the surroundings.

Combustion reactions, such as the complete combustion of methane, are exothermic.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

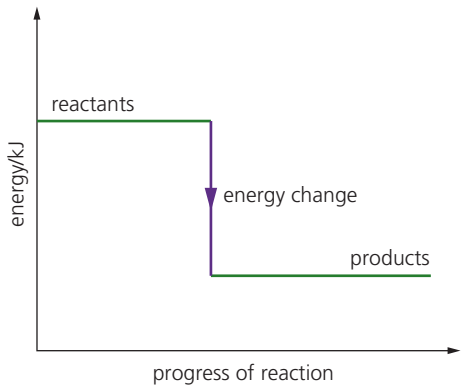
Endothermic reactions are reactions in which thermal energy is taken in from the surroundings.

Thermal decomposition reactions, such as the thermal decomposition of calcium carbonate, are endothermic.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

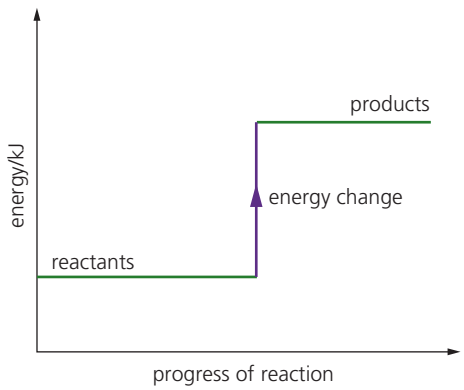
Exothermic and endothermic reactions can be represented by **energy level diagrams**. These diagrams show the energy of the reactants and products, and the energy change as the reaction progresses.

In an exothermic reaction, the products have *less* energy than the reactants (see Figure 6.2). This is because thermal energy is transferred to the surroundings.



▲ **Figure 6.2 Energy level diagram for an exothermic reaction**

In an endothermic reaction, the products have *more* energy than the reactants (see Figure 6.3). This is because thermal energy is taken in from the surroundings.



▲ **Figure 6.3 Energy level diagram for an endothermic reaction**

Revision activity

Make diagrams and graphs more memorable by adding your own extra information. You might use descriptive text (see Figure 1.3, page 3), colours, symbols or drawings to make them stick in your memory. Just remember not to include your additions if you are asked to draw the diagram in an exam.

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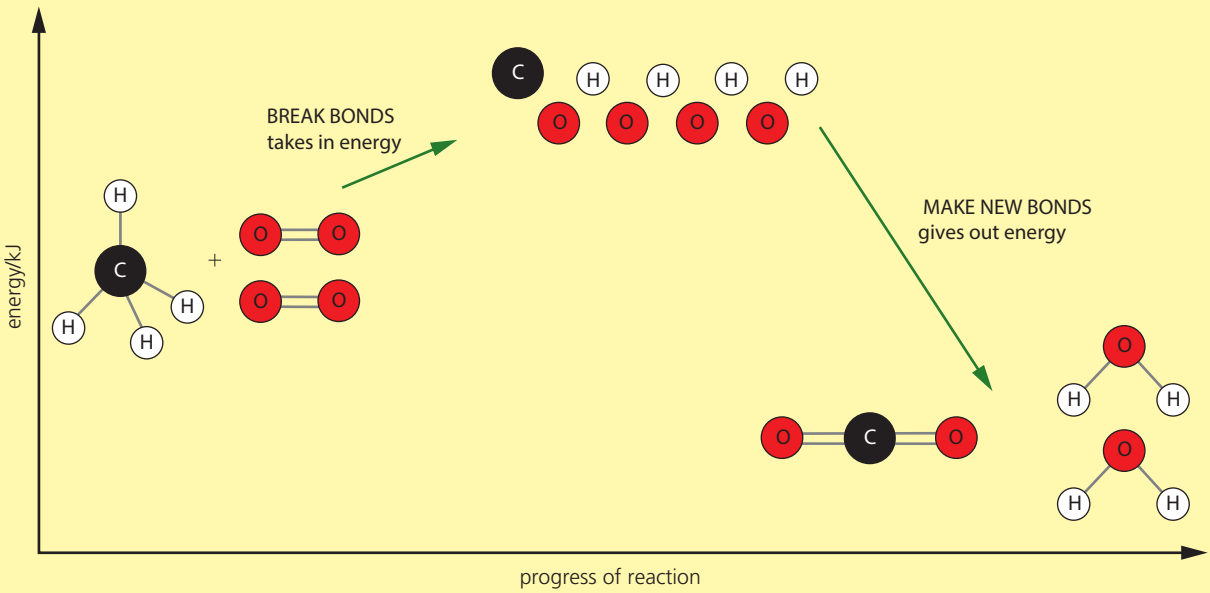
Most chemical reactions involve the breaking of covalent bonds in reactants. When this happens, the molecules change into atoms. The atoms then form new covalent bonds, joining together to form new molecules in the products (see Figure 6.4).

- Breaking of bonds is an endothermic process (energy is taken in).
- Formation of bonds is an exothermic process (energy is given out).

The amount of energy put in to break bonds is very unlikely to be equal to the amount of energy given out when new bonds are formed, so most reactions are either endothermic or exothermic.



▲ **Figure 6.4 Breaking and forming bonds during the combustion of methane**

Bond energies

Bond energy is the amount of energy required to break one mole of covalent bonds in gaseous molecules. It is numerically equal to the amount of energy given out when new bonds form in gaseous molecules.

Enthalpy changes

The transfer of thermal energy during a reaction is called the **enthalpy change**, ΔH.

The enthalpy change is the difference between the thermal energy put in to break the bonds in the reactants and the thermal energy given out when new bonds in the products form.

If less thermal energy is put in to break bonds in the reactants than is given out when new bonds form in the products, the overall reaction is exothermic. ΔH has a negative value.

If more thermal energy is put in to break bonds in the reactants than is given out when new bonds form in the products, the overall reaction is endothermic. ΔH has a positive value.

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Skills

Calculating enthalpy changes

The general equation for bond energy calculations is:

ΔH = energy required to break bonds – energy given out when forming bonds

Worked example

Is the formation of hydrogen chloride from its elements exothermic or endothermic?

Bond energies are shown in Table 6.1.

▼ **Table 6.1 Bond energies**

Bond	Bond energy (kJ/mol)
H–H	435
Cl–Cl	242
H–Cl	432

Answer

The equation can be written to show the structure of the molecules:

$$H–H + Cl–Cl \rightarrow \mathbf{2}H–Cl$$

energy put in to break bonds	**energy given out when bonds form**
H–H = 435 kJ	2 × H–Cl = 2 × 432 = 864 kJ
Cl–Cl = 242 kJ	
total energy put in = 435 + 242 = 677 kJ	total energy given out = 864 kJ

As 864 is a larger number than 677, more energy is given out when the bonds in the products form than has to be put in to break the bonds in the reactants.

Therefore, the reaction is exothermic and the overall energy change is:

$$677 - 864 = -187\,\text{kJ/mol}$$

This means that when 1 mole of gaseous H_2 molecules react with 1 mole of gaseous Cl_2 molecules to form 2 moles of gaseous HCl molecules, 187 kJ of energy are given out to the surroundings.

The enthalpy change for the reaction, ΔH = –187 kJ/mol.

The negative (–) sign indicates that the reaction is exothermic.

Activation energy

Activation energy, E_a, is the minimum amount of energy that particles must contain if they are to react when they collide.

If a flame is applied to a mixture of methane and oxygen, the methane burns rapidly to form carbon dioxide and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

However, if methane is mixed with oxygen, no reaction takes place.

This is because the energy the molecules of methane and oxygen contain is less than the activation energy. The flame provides the molecules with additional energy, so the energy they contain is equal to or greater than the activation energy.

If particles do not contain energy *equal to or greater than* the activation energy, they can still collide with one another, but the collision will not be successful – it will not lead to the production of products.

Energy changes in reactions can be shown on reaction pathway diagrams. In these diagrams:

- a downward arrow ↓ represents an exothermic change, where ΔH is negative
- an upward arrow ↑ represents an endothermic change, where ΔH is positive
- activation energy, E_a, is always positive and is always represented by an upward arrow ↑.

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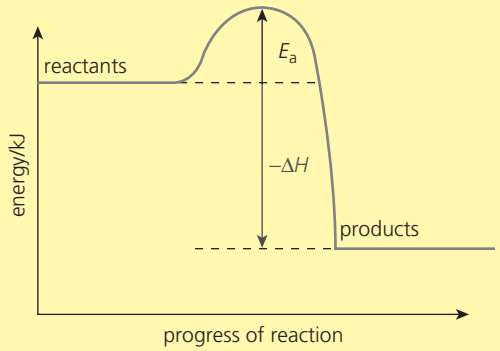
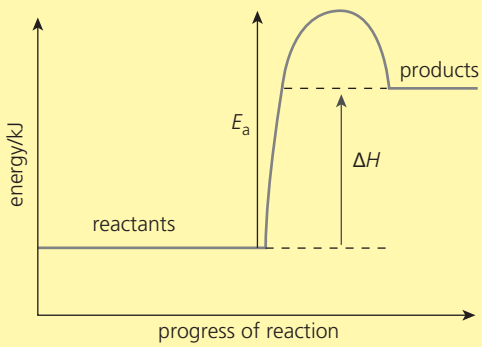


▲ **Figure 6.5 Reaction pathway diagram for an exothermic reaction**



▲ **Figure 6.6 Reaction pathway diagram for an endothermic reaction**

Sample questions

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1 State the differences between an exothermic and an endothermic reaction. [2]

Student's answer

An exothermic reaction transfers heat energy to the surroundings leading to a decrease in the energy of the surroundings. In an exothermic reaction, the reactants have less energy than the products.

An endothermic reaction takes in heat energy from the surroundings leading to an increase in the energy of the surroundings. In an endothermic reaction, the reactants have more energy than the products.

Correct answers

An exothermic reaction transfers thermal energy to the surroundings leading to an increase in the energy, and therefore temperature, of the surroundings. In an exothermic reaction, the reactants have more energy than the products.

An endothermic reaction takes in thermal energy from the surroundings leading to a decrease in the energy, and therefore temperature, of the surroundings. In an endothermic reaction, the reactants have less energy than the products.

Teacher's comments

Exothermic and endothermic reactions are opposites of one another.

The following phrases are used to convey this:

- to and from the surroundings
- decrease and increase in energy
- more and less energy.

The student had the direction of transfer of heat energy correct in both cases. However, the comments concerning increase/decrease in heat energy and more/less energy were both the wrong way round.

2 State and explain the differences between an exothermic and an endothermic reaction in terms of energy changes during bond breaking and bond making.

Student's answer

In an exothermic reaction, the amount of heat energy given out when bonds in the reactants break is less than the amount of energy taken in when bonds in the product are made.

In an endothermic reaction, the amount of heat energy given out when bonds in the reactants break is more than the amount of energy taken in when bonds in the product are made.

Teacher's comments

The student was wrong in stating that:

- energy is given out when bonds break
- energy is taken in when bonds form.

These are both very common incorrect statements.

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Correct answer

In an exothermic reaction, the amount of heat energy required to break the bonds in the reactants is less than the amount of energy given out when bonds in the product are made. Therefore, there is an overall transfer of energy to the surroundings.

In an endothermic reaction, the amount of heat energy required to break the bonds in the reactants is more than the amount of energy given out when bonds in the product are made. Therefore, there is an overall transfer of energy from the surroundings.

Exam-style questions

1 The components of petroleum are separated into fractions by fractional distillation.

a Name the property that the process of fractional distillation depends upon. [1]

b State how the following change from the bottom to the top of the fractionating column:

i viscosity [1]

ii volatility [1]

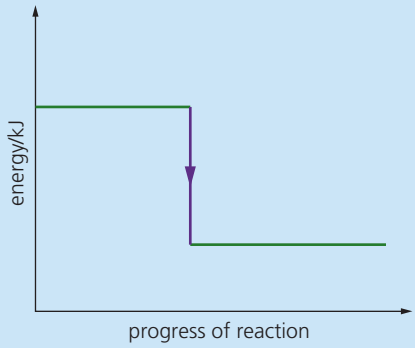
iii chain length [1]

c Complete the table below which shows the uses of different named fractions. [8]

[Total: 12]

Fraction	Use
[1]	Lubricants, waxes or polishes
Refinery gas	[1]
[1]	Making roads
Naphtha	[1]
[1]	Fuel in ships or home heating systems
Gasoline or petrol	[1]
[1]	Fuel for diesel engines
Kerosene or paraffin	[1]

2 The questions which follow are about the energy level diagram opposite.



a Add the words below to label the energy level diagram:

i products [1]

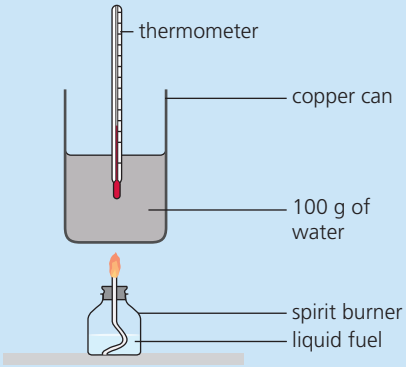
ii energy change [1]

iii reactants [1]

b State whether the reaction is exothermic or endothermic. Explain how you made your decision. [1]

[Total: 4]

3 A student investigated four fuels to find out which gave off the most energy, using the apparatus shown.



a In each experiment, the student used the same amount of fuel.

i Suggest one other factor that should be kept the same in each of the four experiments. [1]

ii The student used the thermometer to stir the water. Suggest why it is important to keep the water stirred. [1]

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The results are shown in the table below.

Fuel	Initial temperature of the water/°C	Final temperature of the water/°C
Ethanol	24	40
Propanol	24	42
Paraffin	22	33
Petroleum spirit	20	40

b Name the fuel that transfers the most energy to the water.
Explain your answer. [2]
[Total: 4]

4 Propane burns in excess oxygen to form carbon dioxide and water according to the equation below.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Calculate the overall energy change occurring when 1 mole of $C_3H_8(g)$ reacts with 5 moles of $O_2(g)$ to form 3 moles of $CO_2(g)$ and 4 moles of $H_2O(g)$ by using the following steps:

a Draw the structures of all the molecules shown in the equation.
Show all the atoms and all the bonds.
(If you have not yet studied Chapter 12, it will help to know that propane has 2 C–C bonds and 8 C–H bonds.) [2]

b Write down the number of moles of each type of bond that have to be broken in the reactants.
(Remember to consider the number of moles of each reactant.) [1]

c Use the values of bond energy from the table below to calculate the total amount of energy that has to be put in to break all the bonds in **(b)**. [1]

Bond	Bond energy/kJ/mol
C–C	347
C–H	435
O=O	497
C=O	803
O–H	464

d Write down the number of moles of each type of bond that have to be formed in the products.
(Remember to consider the number of moles of each product.) [1]

e Calculate the total amount of energy that is given out when all the bonds in **(d)** are formed. [1]

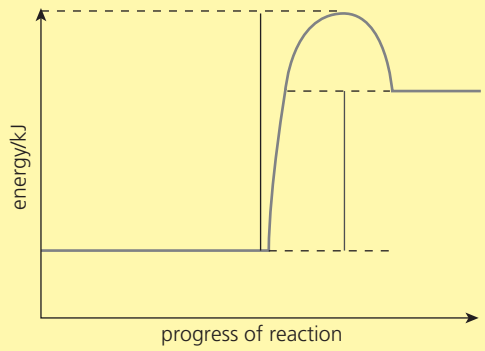
f Use your answers to **(c)** and **(e)** to calculate the overall energy change in the reaction.
State whether the reaction is exothermic or endothermic. [3]

g Write down the value of ΔH for the reaction. Your answer should have a sign and units. [2]
[Total: 13]

5 A reaction pathway diagram is shown below.

a Add the words below to label a copy of the energy level diagram:
- **i** products [1]
- **ii** E_a [1]
- **iii** ΔH [1]
- **iv** reactants [1]

b Use arrow heads to show whether the activation energy and enthalpy change are exothermic or endothermic. [2]
[Total: 6]



Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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7 Chemical reactions

Key objectives

By the end of this section, you should be able to:

Reactions

- identify physical and chemical changes and describe the differences between them

Factors that affect the rate of reaction/enzymes

- describe the effect on the rate of reaction of:
  - changing the concentration of aqueous solutions
  - changing the pressure of gases
  - changing the surface area of solids
  - changing the temperature
  - adding or removing a catalyst, including enzymes
- explain the meaning of the term catalyst
- describe practical methods for investigating the rate of a reaction, including measuring change in mass of a reactant or product and measuring the volume of a gas produced
- interpret data, including graphs, from rate of reaction experiments

- describe collision theory in terms of:
  - number of particles per unit volume
  - frequency of collisions between particles
  - kinetic energy of particles
  - activation energy, E_a
- use collision theory to explain the effect on the rate of reaction of:
  - changing the concentration (of solutions)
  - changing the pressure (of gases)
  - changing the surface area (of solids)
  - changing the temperature
  - adding or removing a catalyst, including enzymes
- evaluate practical methods for investigating the rate of a reaction, including measuring change in mass of a reactant or product and measuring the volume of a gas produced

Reversible reactions and equilibrium

- state that some chemical reactions are reversible and shown using the symbol $\rightleftharpoons$
- describe how changing the conditions can change the direction of the following reversible reactions:
  - the effect of heat on hydrated compounds
  - the addition of water to anhydrous copper(II) sulfate and anhydrous cobalt(II) chloride

- state the factors which indicate that a system is at equilibrium in terms of rates and concentrations
- predict and explain how the position of equilibrium for a reversible reaction is affected by:
  - changing temperature
  - changing pressure (of gases)
  - changing concentration (of solutions)
  - using a catalyst

Ammonia

- state the symbol equation for the production of ammonia in the Haber process
- state the sources of hydrogen and nitrogen in the Haber process
- state the typical conditions in the Haber process

Industrial manufacture of sulfuric acid

- state the symbol equation for the conversion of sulfur dioxide to sulfur trioxide in the Contact process
- state the sources of sulfur dioxide and oxygen in the Contact process
- state the typical conditions for the conversion of sulfur dioxide to sulfur trioxide in the Contact process
- explain the typical conditions used for the Haber process and in the Contact process in terms of:
  - rate of reaction
  - position of equilibrium
  - safety considerations
  - economics

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Key terms

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Term	Definition
Catalyst	A substance which increases the rate of a chemical reaction and is chemically unchanged at the end of the reaction. A catalyst increases the rate of a chemical reaction by providing an alternative reaction path which has a lower activation energy, E_a.
Enzyme	Enzymes are protein molecules which are biological catalysts.
Equilibrium	When a reversible reaction takes place in a closed container and both the forward and reverse reactions occur at the same rate.
Rate of reaction	A measure of the change which happens during a reaction in a single unit of time.
Reversible reaction	A chemical reaction that can go both forwards and backwards. Once some of the products have been formed, they will undergo a chemical change once more to re-form the reactants.

7.1 Reactions

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Physical changes are changes in which new chemical substances are *not* produced. Changes in state, that is melting, boiling, evaporation, condensation and freezing (see Chapter 1), and separation of mixtures, for example filtration, distillation, fractional distillation, chromatography and crystallisation (see Chapter 14), are examples of physical changes.

Chemical changes are changes in which new chemical substances *are* produced (see Section 2.2). Decomposition, electrolysis, respiration, photosynthesis, redox, neutralisation, cracking, addition, substitution, polymerisation and combustion are examples of chemical changes.

Physical properties are the properties of a substance that can be measured and are related to physical changes. Examples are melting point, boiling point and density.

Chemical properties are the properties of a substance that are related to chemical changes. Examples are the things that substances react with and details of such reactions.

A physical property of all metals is that they conduct electricity, whereas a chemical property of some metals is that they react with acids to produce a salt and hydrogen.

7.2 Factors that affect the rate of reaction

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The **rate** of a chemical reaction can be determined by measuring one of the following:

- how the amount of one of the reactants decreases with time
- how the amount of one of the products increases with time

The rate of a reaction can be changed by:

- changing the concentration of a solution
- changing the pressure of a gas
- changing the surface area of a solid
- changing the temperature
- adding a **catalyst**

Rates of reaction are best studied through practical work. You should be able to describe methods to investigate rates of reaction.

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Skills

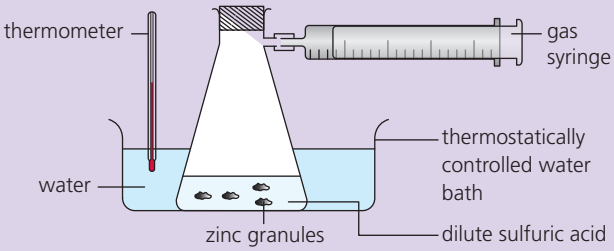
Measuring rate by measuring volume of gas produced

Reactions in which solids react with liquids to produce gases, among other products, are commonly used to investigate rates of reaction. An example is the reaction between zinc, Zn(s), and dilute sulfuric acid, H_2SO_4(aq):

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

Experiment 1

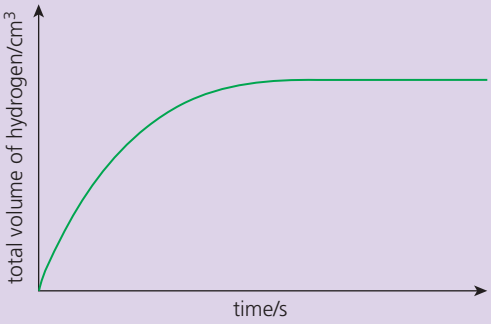
A student used the apparatus shown in Figure 7.1 to investigate the reaction between 50.0 cm^3 of 0.10 mol/dm^3 sulfuric acid and excess zinc granules.



▲ **Figure 7.1 Measuring rate of reaction between zinc and sulfuric acid**

The temperature was kept at 25°C using a water bath. The volume of hydrogen produced was measured at regular time intervals and plotted on a graph (Figure 7.2).

Remember that this type of graph does *not* plot rate against time, but plots mass, concentration or volume of a reactant or product against time.



▲ **Figure 7.2 The volume of hydrogen produced against time for Experiment 1**

The **gradient** shows the rate of reaction. The steeper the gradient, the higher the rate.

In this case:

- the graph is steepest at the start, which means that the rate of reaction is fastest at the start
- the graph then becomes less steep, which means that the rate of the reaction becomes slower
- eventually the graph levels off, which means that no more hydrogen gas is released and the rate of reaction is zero

The rate of any reaction:

- is highest at the start (when $t = 0$) because the concentrations of the reactants are highest at the start
- decreases as time increases because the concentrations of the reactants decrease over time
- becomes zero when one or all of the reactants are used up

Collision theory

In any reaction between gases of the type A(g) + B(g) → C(g), particles of reactants A and B must collide with each other if they are to produce product C. There are two types of collisions:

- successful
- unsuccessful

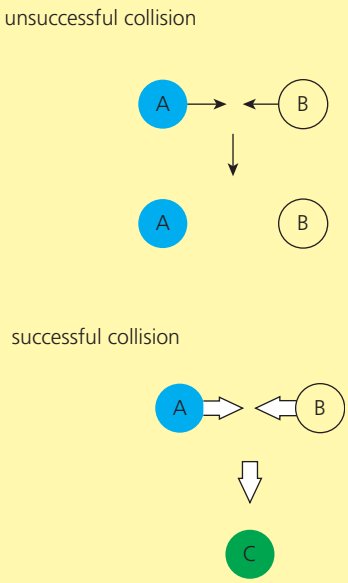
In an unsuccessful collision, particles of A and B merely bounce off each other and remain as A and B.

However, in a successful collision, particles of A and B collide and change into C.

Collisions are only successful if the reacting particles collide with at least a minimum amount of energy called the **activation energy**, E_a.

The rate of a chemical reaction depends on the number of successful collisions in a single unit of time.

If a change is made that increases the number of collisions in a unit of time, the number of successful collisions automatically increases too because a certain proportion of all collisions are always successful.



▲ **Figure 7.3 Collision theory**

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Saying that the rate of a reaction is higher because there are more collisions is an incomplete statement. The correct statement is that the *collision frequency increases*, i.e., there are more collisions in any given amount of time.

In the reaction in Experiment 1, the ionic equation:

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

shows that collisions between zinc atoms and hydrogen ions must take place for the reaction to occur.

- The rate of reaction is fastest at the start because this is when the concentration of hydrogen ions is highest – the number of collisions between hydrogen ions and zinc atoms in any given amount of time is most frequent at the start.
- The rate of reaction then decreases because, as the concentration of hydrogen ions decreases, collisions occur less frequently.
- When all the sulfuric acid is used up, the concentration of hydrogen ions becomes zero. Therefore, there are no more collisions and the rate becomes zero.

While it is possible to refer to the concentration of a gas, it is more usual to consider pressure.

The higher the pressure exerted by a gas, the closer together the molecules and the greater the collision frequency.

Skills

Investigating other factors

The student who carried out Experiment 1 (see page 69) then repeated the investigation, changing a different variable each time, as shown in Table 7.1. The changed variable is shaded.

▼ **Table 7.1 Investigating rates using the reaction of zinc and dilute sulfuric acid**

Experiment	Temperature/°C	Catalyst	Sulfuric acid, $H_2SO_4(aq)$	Zinc, Zn(s)
1	25	None	50.0 cm³ of 0.10 mol/dm³	Granules
2	25	None	25.0 cm³ of 0.20 mol/dm³	Granules
3	25	None	50.0 cm³ of 0.10 mol/dm³	Powder
4	50	None	50.0 cm³ of 0.10 mol/dm³	Granules
5	25	A few drops of aqueous copper(II) sulfate	50.0 cm³ of 0.10 mol/dm³	Granules

Experiment 2: Changing the concentration of aqueous reactant

The concentration of sulfuric acid is doubled in Experiment 2 but the volume of sulfuric acid is halved, which means that the number of moles of sulfuric acid is the same. The graph in Figure 7.4 shows the results from Experiment 2 together with the those from Experiment 1.

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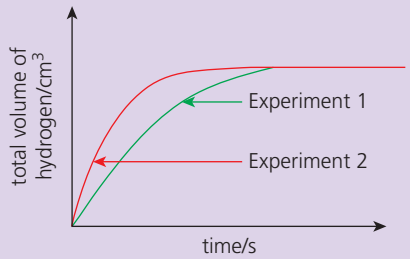


▲ **Figure 7.4 Comparing results for Experiment 1 and Experiment 2**

The graph for Experiment 2 is steeper at the start, which means that the rate of reaction is higher than at the start of Experiment 1.

The graphs level off at the same volume of hydrogen because the amount of hydrogen produced depends on the number of moles of sulfuric acid, which is the same in both experiments (as it is in all five experiments).

Experiment 3: Changing the particle size of solid reactant

When using zinc powder instead of granules, the particle size is decreased. (This is the same as saying the surface area is increased.) The graph of results for this experiment is also steeper at the start than for Experiment 1. This means the rate of reaction is faster.

> Collisions can only occur on the surface of the zinc. With smaller particles, there are more zinc atoms available to collide with the hydrogen ions in any given time. More collisions occurring in any given amount of time means that there are more successful collisions in a unit of time and, therefore, a greater rate of reaction.

Experiment 4: Changing the temperature

At a higher temperature, the graph of volume of hydrogen produced against time is steeper at the start than for Experiment 1. This means that the initial rate of reaction is higher at the higher temperature.

> At a higher temperature, the reacting particles have more kinetic energy. This means that the particles move faster and collision frequency increases. Therefore, the rate of reaction increases.
>
> However, there will also be a greater proportion of collisions where the particles have energy equal to or greater than the activation energy. Therefore, there will be an increase in the frequency of successful collisions. This is the main reason why rates of reaction are faster at higher temperatures.

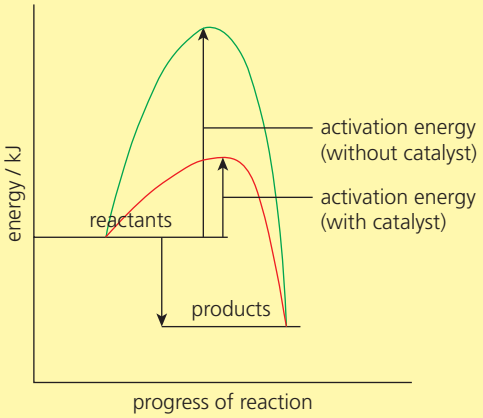
Experiment 5: Using a catalyst

Aqueous copper(II) sulfate acts as a catalyst in this reaction.

When a catalyst is used, the results graph is also steeper at the start than for Experiment 1 – the initial rate of reaction is higher.

Catalysts increase the rate of a reaction and are chemically unchanged at the end of the reaction.

> Catalysts lower the activation energy of a reaction. This means that a greater proportion of collisions have enough energy to be successful collisions. More successful collisions in any given amount of time means the reaction is faster.
>
> The lowering of activation energy in a catalysed reaction can be shown in the reaction pathway diagram below (Figure 7.5).



▲ **Figure 7.5 Catalysts and activation energy**

As can be seen, using a catalyst has no effect on the overall energy change of a reaction, but it lowers the activation energy, thus increasing the rate of reaction.

As catalysts are unchanged at the end of a reaction, it is easy to think they do not take part in the reaction. This is not the case – the increasing rate suggests that catalysts have a considerable part to play.

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Revision activity

Make a large copy of Figure 7.4 in the middle of a sheet of paper. Use different colours to add curves for Experiments 3, 4 and 5. Add notes in the respective colours around the edge of the graph to explain the similarities and differences between the curves.

7.3 Enzymes

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Enzymes are protein molecules that act as biological catalysts. (The term *biocatalyst* means something slightly different so always write *biological catalyst* in full.)

Reactions catalysed by enzymes are affected by the same factors as reactions that use non-biological catalysts.

An important exception to this is temperature. The rate of a reaction that is catalysed by enzymes increases as the temperature increases only up to a certain point. Above this temperature, the rate decreases because the structure of the enzyme is altered and it loses its ability to catalyse the reaction. We say that the enzyme is **denatured**.

The temperature at which an enzyme causes the maximum rate of reaction is called the **optimum temperature**.

7.4 Reversible reactions and equilibrium

REVISED

Reversible reactions

Some reactions can be reversed by changing the conditions.

Skills

Hydrated and anhydrous compounds

If crystals of hydrated copper(II) sulfate and hydrated cobalt(II) chloride are heated, they change colour as they lose their water of crystallisation and become anhydrous salts.

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$

blue crystals → white powder

$$CoCl_2.6H_2O(s) \rightarrow CoCl_2(s) + 6H_2O(g)$$

pink crystals → blue powder

However, in both cases, the reactions can be made to proceed in the reverse direction by adding water to the anhydrous salts, in which case the crystals form again, as can be seen by the reverse colour change.

$$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_2O(s)$$

white powder → blue crystals

$$CoCl_2(s) + 6H_2O(l) \rightarrow CoCl_2.6H_2O(s)$$

blue powder → pink crystals

These reactions are called **reversible reactions**. They can be made to proceed in the reverse direction by changing the conditions.

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Equilibrium

If a reversible reaction is allowed to proceed in a closed container, it reaches a state that is known as **chemical equilibrium**.

If a mixture of hydrogen and iodine gases is heated in a closed container, the hydrogen reacts with the iodine to produce hydrogen iodide:

$$H_2(g) + I_2(g) \rightarrow 2HI(g)$$

This is called the **forward reaction**.

As soon as hydrogen iodide molecules are formed, they start to decompose into hydrogen and iodine:

$$2HI(g) \rightarrow H_2(g) + I_2(g)$$

This is called the **reverse** (or backward) **reaction**.

Therefore, two reactions are occurring in the same container at the same time.

One reaction is the reverse of the other. This can be shown by the following expression:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

$$\text{reactants} \rightleftharpoons \text{products}$$

The forward reaction starts off quickly and the rate decreases as the concentrations of hydrogen and iodine decrease.

The backward reaction starts off slowly and the rate increases as the concentration of hydrogen iodide increases.

Eventually, both rates become *equal*. The system is then in a state of chemical equilibrium. At this point, the reactants and products are being used up and produced at the same rate. Therefore, their concentrations are no longer changing and become constant.

If you are asked to describe the characteristics of an equilibrium system in an exam, do not make any of the following common errors.

- The forward reaction is equal to the reverse reaction.
  This is a meaningless statement unless the word *rate* is used.
- The amounts of reactants and products no longer change.
  In this case, the word *amounts* must be replaced by *concentrations*.
- The concentrations of products and reactants become equal.
  This is incorrect – the concentrations of products and reactants no longer change, but the actual concentration of the reactants may be higher than that of the products (or the other way around).

Characteristics of equilibrium systems

Equilibrium can only occur in a closed system (closed container), in which no substances can escape or enter from the outside.

- The rate of the forward reaction is equal to the rate of the reverse reaction.
- The concentrations of all reactants and products become constant.

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Effects of changing the conditions of an equilibrium system

▼ **Table 7.2 How changes affect equilibrium position**

Change	Effect on equilibrium position
Increase temperature	Shifts in the endothermic direction
Increase pressure of gases	Shifts to form fewer gas molecules
Increase concentration of reactants in solution	Shifts to form more products
Add catalyst	No change

Decreases in concentration, pressure and temperature have the opposite effect to increases. For example, the following equation represents an equilibrium:

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

The forward reaction is exothermic.

This means that $2NO_2(g) \rightarrow N_2O_4(g)$ is an exothermic reaction.

Therefore, $N_2O_4(g) \rightarrow 2NO_2(g)$ is an endothermic reaction.

The equation shows that there are two gas molecules on the left-hand side of the equilibrium sign and one gas molecule on the right-hand side of the equilibrium sign.

▼ **Table 7.3 How changes in conditions affect the reaction $2NO_2(g) \rightleftharpoons N_2O_4(g)$**

Change	Effect on equilibrium position	Result in this example
Increase the concentration of reactants (NO_2)	Shifts to the right (in the direction of products)	Concentration of products (N_2O_4) increases
Increase the concentration of products (N_2O_4)	Shifts to the left (in the direction of reactants)	Concentration of reactants (NO_2) increases
Increase the total pressure	Shifts in the direction of fewer molecules	Concentration of products (N_2O_4) increases
Increase temperature	Shifts in the endothermic direction	Concentration of reactants (NO_2) increases
Add a catalyst	Increases the rate of both forward and reverse reactions, but does not change the position of the equilibrium	No change

Decreases in concentrations, pressure and temperature have the opposite effect to increases.

Here are some more common mistakes which students make when they are asked questions about equilibrium reactions.

- The equilibrium shifts to the exothermic side.

  There is no *exothermic side*. If the question is about the reaction above, you should say 'the equilibrium shifts in the direction of the forward reaction'.
- The equilibrium shifts towards the reaction with fewer molecules.

  There is no reaction with fewer molecules. A correct statement would be 'the equilibrium shifts in the direction of fewer molecules'.

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7.5 Ammonia – an important nitrogen-containing chemical

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Ammonia has many industrial uses. It is manufactured from nitrogen and hydrogen in the Haber process.

- Nitrogen is obtained from the fractional distillation of liquid air.
- Hydrogen is obtained from methane.

Nitrogen and hydrogen react to produce ammonia in a reversible reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

The forward reaction is exothermic.

The gases are:

- passed over a catalyst of iron
- at a temperature of 450°C
- at a pressure of 200 atmospheres/20 000 kPa

The mixture that comes out of the reaction chamber contains about 15% ammonia. The ammonia is liquefied to separate it from the unreacted nitrogen and hydrogen, which are fed back over the catalyst again. Eventually, all the nitrogen and hydrogen are converted into ammonia.

7.6 Industrial manufacture of sulfuric acid – the Contact process

REVISED

Sulfuric acid, which has many industrial uses, is made from sulfur trioxide. The Contact process makes sulfur trioxide from sulfur dioxide and oxygen.

- Sulfur dioxide is obtained from burning sulfur in air or roasting sulfide ores in air.
- Oxygen is obtained from the fractional distillation of liquid air.

Sulfur dioxide and oxygen react together in a reversible reaction to produce sulfur trioxide:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

The forward reaction is exothermic.

The gases are:

- passed over a catalyst of vanadium(V) oxide
- at a temperature of 450°C
- at a pressure of 2 atmospheres/200 kPa

The mixture that comes out of the reaction chamber contains unreacted sulfur dioxide and oxygen as well as sulfur trioxide. The sulfur trioxide is separated from the unreacted sulfur dioxide and oxygen, which are passed over the catalyst again. Eventually, all the sulfur dioxide and oxygen are converted into sulfur trioxide.

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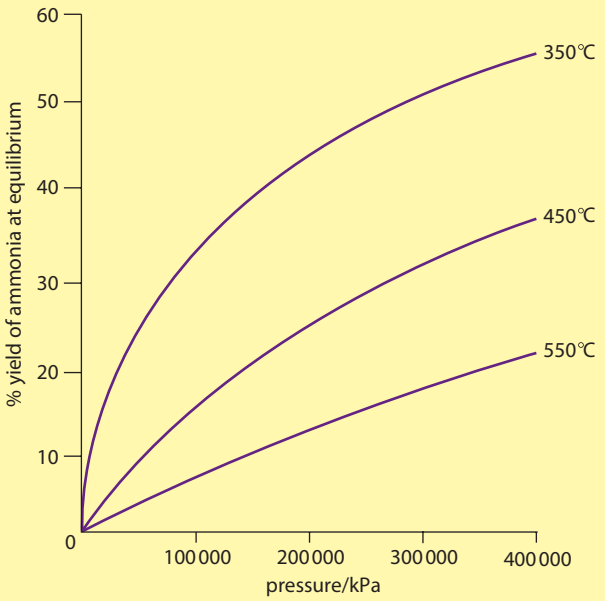
Reasons for conditions

The Haber process and the Contact process are carried out in order to:

- produce the maximum yield of product (equilibrium towards the product side)
- produce the product as quickly as possible (high rates)
- maximise profits (economics)
- minimise hazards (safety)

▼ **Table 7.4 Conditions in the Haber and Contact processes**

Change	Effect of change
Lower temperature	Rate decreases because rates decrease as temperatures decreases
Higher temperature	Yield of product decreases because equilibrium shifts in the endothermic direction, which is to the left
Lower pressure	Rate decreases because rates decrease as pressure of gases decreases Yield of product decreases because the equilibrium shifts in the direction of more molecules
Higher pressure	Rate increases because rates increase as pressure of gases increases Yield of product increases because the equilibrium shifts in the direction of fewer molecules
No catalyst	Rate decreases No effect on yield of product



▲ **Figure 7.6 Yields from the Haber process**

It is not possible to change the temperature in a way that improves both yield and rate, so a temperature of 450°C is a compromise temperature for both processes.

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A higher pressure would improve both the yield and rate but:

- containers made of steel thick enough to withstand higher pressures are expensive, so this would lead to lower profits
- it would be hazardous because it increases the risk of gas leaks and explosions

Since the yield and rate are both satisfactory when using the stated pressures, the additional cost and risks of using higher pressures are uneconomic.

Without using a catalyst, the rate would decrease.

Revision activity

Make flow charts to show the steps in the Haber process and the Contact process. Remember to show how the reactants are supplied to the reaction chamber and add loops to show how unreacted gases are reused.

Sample questions

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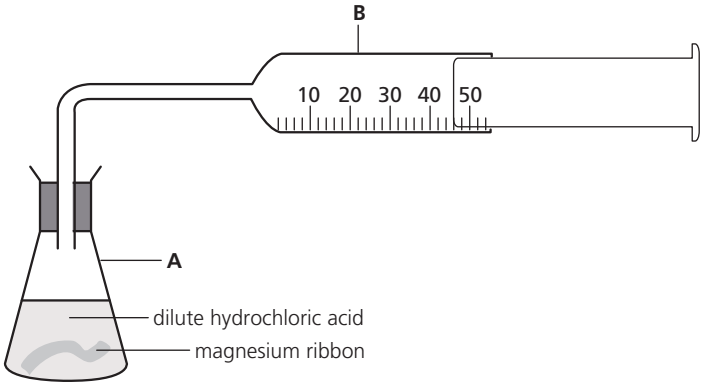
1 Magnesium was added to excess dilute hydrochloric acid in apparatus **A** below.



A stop clock was started at the same time.
Hydrogen gas was collected in apparatus **B**. The volume of gas was measured at regular time intervals.

a Name:
- **i** apparatus **A** [1]
- **ii** apparatus **B** [1]

b Name a piece of apparatus that could be used instead of **B** to collect the hydrogen gas and measure its volume. [1]

c State the volume of hydrogen gas collected in apparatus **B**. [1]

d Give a test for hydrogen gas. State the result of the test. [1]

e The reaction stops after 30 minutes. State why the reaction stops. [1]

f The rate of the reaction is fastest at the start. Explain why. [1]

g The rate of the reaction decreases as the time increases. Explain why. [2]

h State one improvement that could be made to the apparatus to make sure that the temperature does not change during the reaction. [1]

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Student's answers

a i flask
ii syringe
b test-tube
c 43 cm^3
d glowing splint pops
e The reactants have been used up.
f The most acid is present at the start.
g The concentration of acid decreases.
h Use a water bath.

Teacher's comments

a i *Flask* is too general a term. There are many types of flasks.
ii **B** should be described as a *gas syringe*.
b A test-tube can be used to collect the gas, but not to measure the volume.
c The student counted the number of divisions but failed to realise that each division represents 2 cm^3.
d A glowing splint is used to test for oxygen.
e The question states that the dilute hydrochloric acid is in excess. Therefore, the student should have realised that the magnesium is used up. Using the term *reactants* is not specific enough.
f Students are supposed to know that the rate of a reaction depends on the concentration of an aqueous solution. Therefore, the word *concentration* should have been used. Another common error is to say the concentration is high at the start rather than at its highest.
g The student's answer was correct.
h A water bath is the correct piece of apparatus, but it should be clear that the water is kept at a constant temperature, for example by using a thermostat.

Correct answers

a i conical flask
ii gas syringe
b inverted burette (or measuring cylinder) containing water
c 46 cm^3
d A lighted splint pops.
e All the magnesium is used up.
f The concentration of the hydrochloric acid is highest at the start.
g The concentration of the hydrochloric acid decreases.
h thermostatically controlled water bath

2 The forward reaction in the Haber process is exothermic.
What happens to the position of equilibrium when the temperature increases?

Student's answer

The rate of the reverse reaction increases because the forward reaction is exothermic.

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Teacher's comments

It is helpful to treat equilibrium and rate as two completely separate topics.

The student should not have used the word *rate*. If temperature of an equilibrium system is increased, the rate of both forward and reverse reactions is increased. An increase in temperature speeds up all reactions except those catalysed by enzymes that are already at or above the optimum temperature.

Correct answers

The equilibrium shifts in the endothermic direction – to the left.

OR

The equilibrium shifts to the left because the forward reaction is exothermic.

Exam-style questions

1 Explain whether the following are chemical changes or physical changes:

- **a** dissolving sodium chloride in water [1]
- **b** electrolysis of aqueous sodium chloride [1]
- **c** cracking alkanes [1]
- **d** fractional distillation of liquid air [1]
- **e** separating the dyes in ink by chromatography [1]

[Total: 5]

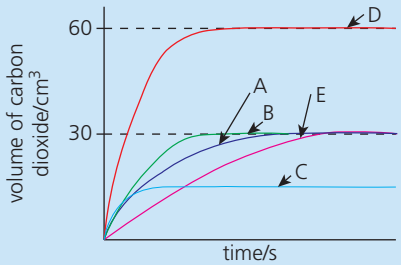
2 When an excess of marble chips (calcium carbonate) is added to 50 cm³ of 0.10 mol/dm³ hydrochloric acid at 25°C, the following reaction occurs:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

The volume of carbon dioxide gas was collected in a gas syringe and measured at regular time intervals. This was Experiment 1.
The experiment was repeated as shown in the table below. The calcium carbonate is in excess in all five experiments.

Experiment	Hydrochloric acid	Calcium carbonate	Temperature/°C	Graph
1	50 cm³ of 0.10 mol/dm³	Marble chips	25	A
2	50 cm³ of 0.20 mol/dm³	Marble chips	25	[1]
3	50 cm³ of 0.10 mol/dm³	Powdered	25	[1]
4	50 cm³ of 0.10 mol/dm³	Marble chips	12.5	[1]
5	50 cm³ of 0.10 mol/dm³	Marble chips	50	[1]

The graphs plotted in each case are shown below.



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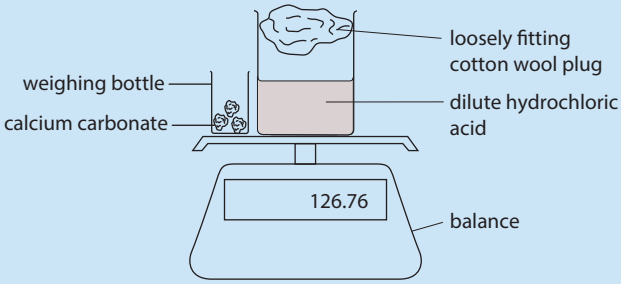
Add letters to the table to show which graph corresponds to each experiment.
Each letter may be used once, more than once or not at all. [Total: 4]

3 Calcium carbonate reacts with dilute hydrochloric acid:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Bubbling is seen as carbon dioxide gas is given off.
A student investigates the rate of this reaction using samples of calcium carbonate. Each sample has a different particle size.



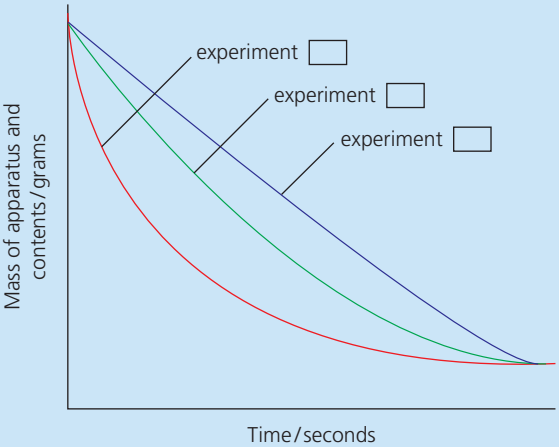
In each experiment, the student adds an excess of calcium carbonate to the dilute hydrochloric acid in the beaker. The weighing bottle is replaced on the balance.
In Experiment 1, the student uses large lumps of calcium carbonate.

a Name the variable, other than mass, that is measured in this experiment. Name the piece of apparatus used to measure this variable. [2]

b State why the mass of the beaker and its contents decrease during the experiment. [1]

c The student does two more experiments.
In Experiment 2, the student uses small lumps of calcium carbonate.
In Experiment 3, the student uses powdered calcium carbonate.
The calcium carbonate is in excess in all three experiments.
Suggest two variables that should be kept constant so that the particle size of the calcium carbonate is the only variable which affects the rate of reaction. [2]

d The student plots graphs of all the results.



i Describe how the graphs are used to decide which experiment has the greatest rate. [1]

ii Write a number in each box on the graph to identify Experiments 1, 2 and 3. [1]

iii State how the graphs show that the reaction stops. [1]

iv State why the reaction stops. [1]

[Total: 9]

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4 Nitrogen and oxygen (both from the air) react in car engines to produce nitrogen dioxide. The equation is:

$$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$$

Use your knowledge of collision theory to explain why the rate of this reaction is faster as the temperature increases. [Total: 3]

5 State in which direction (if any) each of the following equilibrium mixtures would shift if the pressure on the system was increased. Explain your answer in each case.

a $H_2(g) + I_2(s) \rightleftharpoons 2HI(g)$ [1]

b $2O_3(g) \rightleftharpoons 3O_2(g)$ [1]

c $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$ [1]

[Total: 3]

6 State in which direction (if any) each of the following equilibrium mixtures would shift if the temperature on the system was decreased. Explain your answer in each case.

a $H_2(g) + I_2(s) \rightleftharpoons 2HI(g)$
exothermic in the forward direction [1]

b $CH_3COOH(g) + H_2O(g) \rightleftharpoons CH_3OH(g) + CO_2(g) + H_2(g)$
endothermic in the forward direction [1]

[Total: 2]

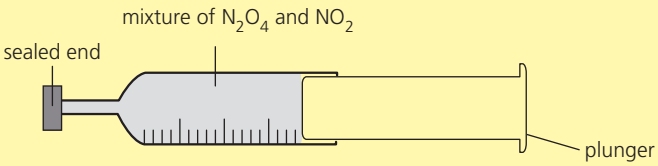
7 Dinitrogen tetroxide, N_2O_4, decomposes into nitrogen dioxide, NO_2. The reaction is reversible.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

$N_2O_4(g)$ is colourless.

$NO_2(g)$ is brown.

A gas syringe containing a mixture of $NO_2(g)$ and $N_2O_4(g)$ was sealed and heated.
After reaching equilibrium, the mixture was a pale brown colour.



a State what is meant by the term *equilibrium*. [2]

b The plunger of the gas syringe is pushed in. The temperature does not change. The mixture initially turns darker brown. After a few seconds, the mixture turns lighter brown because the equilibrium shifts to the left.

i Explain why the mixture initially turned darker brown. [1]

ii Explain why the position of equilibrium shifts to the left. [1]

c The forward reaction is endothermic.

i State what happens to the position of equilibrium when the temperature of the mixture is increased. [1]

ii State what happens to the rate of the forward reaction and the rate of the backward reaction when the equilibrium mixture in the syringe is heated. [2]

[Total: 7]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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8 Acids, bases and salts

Key objectives

By the end of this section, you should be able to:

- describe the characteristic reactions of acids
- state that bases are oxides or hydroxides of metals
- state that alkalis are soluble bases
- describe the characteristic reactions of bases
- state that aqueous solutions of acids contain H^+ ions and aqueous solutions of alkalis contain OH^- ions
- describe how to use pH, as measured with universal indicator paper, to compare hydrogen ion concentration, neutrality, relative acidity and alkalinity
- describe the neutralisation reaction

- define acids and bases in terms of proton transfer
- explain the difference between weak and strong acids in terms of dissociation

- classify oxides as acidic, basic or amphoteric, with examples
- describe the general solubility rules for salts
- describe the preparation of soluble salts by the reaction of an acid with:
  - an alkali
  - excess metal, insoluble base or carbonate

- describe the preparation of insoluble salts by precipitation

- define a hydrated substance and an anhydrous substance

- define the term water of crystallisation in crystals, including $CuSO_4.5H_2O$ and $CoCl_2.6H_2O$

Key terms

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Term	Definition
Acid	A substance which dissolves in water, producing $H^+(aq)$ ions as the only positive ion. A proton (H^+) donor.
Alkali	A soluble base which produces OH^- (aq) ions in water.
Anhydrous salt	A salt which has had its water of crystallisation removed.
Base	A substance which neutralises an acid, producing a salt and water as the only products. Bases are oxides or hydroxides of metals. (Ammonia is also a base.) A proton (H^+) acceptor.
Indicator	A substance that shows whether a substance is acidic or alkaline by changing colour.
Neutralisation	The process in which an acid reacts with a base to form water.
pH scale	A scale running from 0 to 14 used to express the acidity or alkalinity of a substance.
Saturated solution	A solution containing the maximum amount of dissolved solute in the solvent at a given temperature.
Water of crystallisation	Water incorporated into the structure of a substance as it crystallises, for example in copper(II) sulfate pentahydrate ($CuSO_4.5H_2O$).

8.1 Acids and alkalis

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Acids

Acids are substances that produce H^+ ions when they are dissolved in water.

Acids are defined as proton (H^+) donors.

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Bases and alkalis

Bases that do not dissolve in water are known as insoluble bases.

Alkalis are bases that dissolve in water.

Alkalis are substances that produce OH^- ions when dissolved in water.

The two most common laboratory alkalis are aqueous sodium hydroxide and potassium hydroxide. They both exist completely as ions in aqueous solution.

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$
$$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$$

An aqueous solution of ammonia is a base. An aqueous solution of ammonia exists mainly as NH_3 molecules, a small number of which react with water molecules to produce ions.

$$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

Bases are defined as proton (H^+) acceptors.

NH_3 accepts H^+ from H_2O, forming NH_4^+. Thus, NH_3 is acting as a base.

Indicators

Litmus, thymolphthalein and methyl orange can be used as **indicators** to show whether substances are acids or alkalis, but give no information about acid strength.

▼ **Table 8.1 Indicators**

	Litmus	Thymolphthalein	Methyl orange
Colour in acidic solution	Red	Colourless	Red
Colour in neutral solution	Purple	Pale blue	Orange
Colour in alkaline solution	Blue	Blue	Yellow

Neutralisation

All aqueous solutions of acids contain $H^+(aq)$.

All aqueous solutions of alkalis contain $OH^-(aq)$.

When an acid and an alkali react with one another, the $H^+(aq)$ ions in the acid **neutralise** the $OH^-(aq)$ in the alkali, and $H_2O(l)$ is the product.

Skills

Writing ionic equations

You can write ionic equations for any reaction by following these steps:

1. Start from a balanced equation with state symbols.
2. Anything with (aq) as a state symbol should be written as ions.
   - a A dilute acid, e.g. $HCl(aq)$, is written as $H^+(aq)$ and $Cl^-(aq)$.
   - b A metallic compound, e.g. $CuSO_4(aq)$, is written as $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$.
   - c An ammonium salt, e.g. $(NH_4)_2SO_4(aq)$, is written as $2NH_4^+(aq)$ and $SO_4^{2-}(aq)$.
3. Numbers in front of formulae in equations mean that everything *after* the number is multiplied, e.g. $2HNO_3(aq)$ is written as $2H^+(aq)$ and $2NO_3^-(aq)$.

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4 The formulae of any substances with state symbols (s), (*l*) or (g) are not written as ions, thus are not changed in an ionic equation.
5 Any ions which are the same on both sides, known as **spectator ions**, are cancelled.

Worked example

Write an ionic equation for the reaction:

hydrochloric acid + sodium hydroxide → sodium chloride + water

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Answer

HC*l*(aq), NaOH(aq) and NaC*l*(aq) can be written as ions.

$H_2O(l)$ exists as molecules.

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

$Na^+(aq)$ and $Cl^-(aq)$ are present on both sides of the equation. They are spectator ions because they are not changed in the reaction. Therefore, they can be crossed out.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$$

Therefore, the final ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

This is the ionic equation for the reaction between any dilute acid and any aqueous alkali.

For a reaction with ammonia, for example:

$$2NH_3(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$$

the ionic equation is:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$$

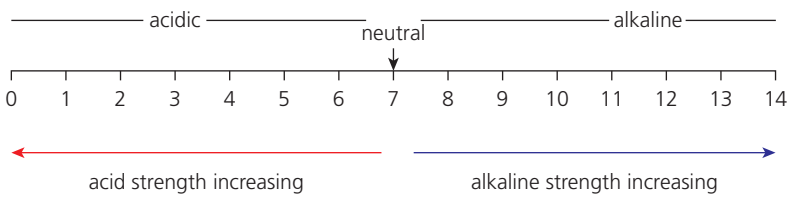
Revision activity

Use highlighters or coloured pens or pencils to make key information about indicators, acids and bases stand out. Do this in your own notes and in books you use – as long as no one else will be using the book after you, of course.

Strong and weak acids and alkalis

Strong and weak acids can be distinguished experimentally using universal indicator paper.

Figure 8.1 shows the **pH scale**, which uses numbers to distinguish between acids and alkalis of different strengths.



▲ **Figure 8.1 The pH scale**

The lower the pH number, the stronger the acid. The higher the pH number, the stronger the alkali.

Strong acids are regarded as having a pH of 0–2. Strong alkalis are regarded as having a pH of 12–14.

Universal indicator shows approximate pH numbers by changing colour. as shown in Table 8.2.

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▼ **Table 8.2 Universal indicator**

Approximate pH	Colour of universal indicator paper
Less than 3	Red
3–6	Orange–yellow
7	Green
8–11	Blue
More than 11	Purple

If a strong and a weak acid of the same concentration are compared, the strong acid contains a higher concentration of $H^+(aq)$ ions than the weak acid.

The common laboratory strong acids are dilute hydrochloric acid, HCl, dilute nitric acid, HNO_3, and dilute sulfuric acid, H_2SO_4.

An aqueous solution of a strong acid does not contain any molecules – they exist completely as ions. For example:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

The $\rightarrow$ in the equation shows that strong acids completely dissociate in aqueous solution.

In aqueous solutions, weak acids, such as ethanoic acid, CH_3COOH, exist mainly as covalent molecules. Only a small number of the molecules dissociate into ions. For example:

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

The equation contains $\rightleftharpoons$ to show partial dissociation. Weak acids partially dissociate in aqueous solution.

8.2 Formation of salts

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Salts are ionic substances formed when the positive hydrogen ions in an acid are replaced by positive metallic ions or ammonium ions.

Solubility rules

Only some salts are soluble in water.

▼ **Table 8.3 Solubility of salts**

Soluble	Partially soluble	Insoluble
All nitrates		
All sodium, potassium and ammonium salts		
Most chlorides		Lead and silver chlorides
Many sulfates		Lead, calcium and barium sulfates
Lead nitrate		All other lead salts
Sodium, potassium and ammonium carbonates		All other carbonates
Sodium and potassium hydroxides	Calcium hydroxide	All other hydroxides

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Oxides

Oxides can be put into three categories.

- **Acidic oxides** are non-metallic oxides that neutralise alkalis and form salts. Examples are carbon dioxide, CO_2, and sulfur dioxide, SO_2. These oxides all dissolve in water and react with water to form acids.
- **Basic oxides** are metallic oxides that neutralise acids and form salts. Examples are calcium oxide, CaO, and copper(II) oxide, CuO. Some basic oxides dissolve in water to form alkaline hydroxides, whereas others are insoluble in water.

- Some metallic oxides are **amphoteric** oxides, which means they react with both acids and bases to form a salt and water. Examples are zinc oxide, ZnO, and aluminium oxide, Al_2O_3.

8.3 Methods of preparing soluble salts

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Using acids to prepare salts:

- hydrochloric acid, HCl, is used to prepare chlorides
- nitric acid, HNO_3, is used to prepare nitrates
- sulfuric acid, H_2SO_4, is used to prepare sulfates (or hydrogen sulfates)

The dilute acids can be reacted with:

a excess metal
b excess insoluble base
c excess insoluble carbonate
d alkali (soluble base) by titration

The positive ion in the salt comes from the metal, insoluble base or carbonate, or alkali.

(a) Acid + metal

Acids react with metals above hydrogen in the reactivity series (although it would be dangerous to use a Group I metal or anything below calcium in Group II in a reaction with acids). The general equation is:

acid + metal → salt + hydrogen

The solid metal disappears, bubbles are seen and a solution of the salt forms. The colour of the solution depends on the metal used. An example is:

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

(b) Acid + base

With insoluble bases, the solid dissolves and a solution forms. No bubbles are seen because no gas is produced. An example is:

$$Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$

Acids react with bases to form a salt and water although, in the case of ammonia, an ammonium salt is the only product.

The general equations are:

acid + base → salt + water

acid + ammonia → ammonium salt

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(c) Acid + carbonate

Acids react with carbonates. The general equation is:

acid + carbonate → salt + water + carbon dioxide

The carbonate may be solid or (if soluble) in solution.

The solid carbonates disappear. In both cases, bubbles are seen and an aqueous solution of the salt forms. The colour of the solution depends on the carbonate used. An example is:

$$CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

(d) Acid + alkali

When dilute acids are added to alkalis, there are no observations (unless an indicator is present) as a colourless solution is produced from two colourless solutions.

An example of this reaction is:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Reactions (a), (b) and (c) use Method 1 below. Method 2, titration, is usually used for reaction (d).

Skills

Preparing salts in the laboratory – Method 1

A solid metal, metal oxide, metal hydroxide or metal carbonate is added to a dilute acid until an excess of the solid is present. The excess solid is removed by filtration and crystals are made from the filtrate by crystallisation and drying.

Worked example

Describe how to prepare a sample of copper(II) sulfate in the laboratory.

Answer

Copper metal cannot be used because it is lower than hydrogen in the reactivity series. Copper(II) oxide, copper(II) hydroxide or copper(II) carbonate could be used instead.

Using copper(II) oxide, the equation is:

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

- Use a spatula to add solid copper(II) oxide (or hydroxide or carbonate) to dilute sulfuric acid in a beaker.
- Stir and/or heat the mixture.
- Continue adding the solid, while stirring, until it will no longer dissolve. This means that all the acid has reacted and the solid is in excess. (If copper(II) carbonate is used, there will be no further bubbling when all the acid has reacted.) The undissolved solid will be visible.
- Filter off the excess solid.
- Make pure crystals of copper(II) sulfate by crystallisation, washing and drying (see Chapter 14).

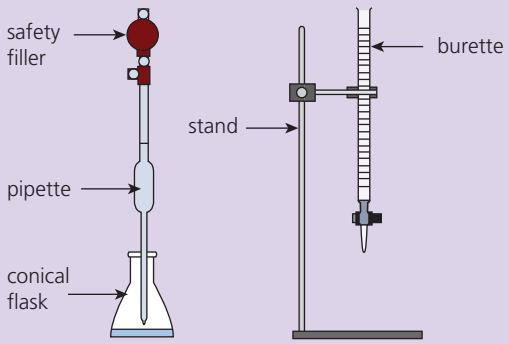
Skills

Preparing salts in the laboratory – Method 2: Titration

Titration uses the equipment shown in Figure 8.2 to determine the volumes of two aqueous solutions that react with each other – neither is in excess. The exact volumes of the two solutions are then mixed and the salt is obtained by crystallisation and drying.

Worked example

Describe how to obtain a sample of sodium sulfate crystals by titrating a suitable acid and alkali.



▲ **Figure 8.2 Equipment for titration**

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Answer

Sodium sulfate crystals can be made with sodium hydroxide and sulfuric acid. The equation is:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

- Use a pipette to transfer 25.0 cm^3 of aqueous sodium hydroxide, NaOH, into a conical flask.
- Add 2–3 drops of methyl orange or thymolphthalein indicator.
- Fill a burette with dilute sulfuric acid, H_2SO_4.
- Add the H_2SO_4 from the burette to the conical flask, approximately 1 cm^3 at a time. Swirl the contents of the flask after each addition.
- Continue until the end point – when the indicator changes colour. This gives an approximate value of the volume of H_2SO_4 required to neutralise the NaOH.
- Carry out an accurate titration. When the end point is close, add the H_2SO_4 one drop at a time. Swirl after each addition. Proceed until the indicator changes colour.
- Carry out more accurate titrations until two volumes are within 0.10 cm^3 of each other.
- Repeat the process without indicator, but using the same volume of acid and alkali as used in the titration.
- Make pure crystals of sodium sulfate by crystallisation and drying (see Chapter 14).

Revision activity

Flow charts are a good way to show processes that have several steps. Create flow charts of your own for the methods of preparing salts described in this chapter.

Reactions of bases

As described above, bases neutralise acids.

Insoluble bases and alkalis react when heated with ammonium salts, giving off ammonia gas. The general equation is:

base + ammonium salt → salt + ammonia + water

For example:

$$Ca(OH)_2(s) + 2NH_4Cl(s) \rightarrow CaCl_2(s) + 2NH_3(g) + 2H_2O(l)$$

8.4 Preparing insoluble salts

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Insoluble salts are made by **precipitation**. This involves mixing two aqueous solutions. The insoluble solid forms as a precipitate, which can be separated and purified.

Lead sulfate can be made by this method.

- As lead nitrate is the only soluble lead salt, aqueous lead nitrate must be used. It can be mixed with any solution that contains aqueous sulfate ions – dilute sulfuric acid or a solution of any soluble sulfate, such as sodium sulfate.
- The precipitate of lead sulfate can be filtered out, washed with distilled water and dried in a low oven, between filter papers or on a warm windowsill.
- The equation is:

$$Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$$

An ionic equation for any precipitation reaction always shows the two aqueous ions on the left and the solid precipitate on the right. In this case:

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$$

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8.5 Testing for different salts

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Tests for **anions** (negative ions) are shown in Table 8.4.

▼ **Table 8.4 Testing for anions**

Test	Result	Anion
Add dilute nitric acid, followed by aqueous silver nitrate	White precipitate	Chloride, Cl^-
	Cream precipitate	Bromide, Br^-
	Yellow precipitate	Iodide, I^-
Add any dilute acid	Bubbles Gas given off turns limewater milky (gas is CO_2)	Carbonate, CO_3^{2-}
Add dilute nitric acid, followed by aqueous barium nitrate	White precipitate	Sulfate, SO_4^{2-}
Add aqueous sodium hydroxide, followed by aluminium; warm gently	Gas given off turns damp red litmus paper blue (gas is NH_3)	Nitrate, NO_3^-

Revision activity

Combine information from Tables 8.3 and 8.4 into a poster or infographic that would help someone trying to identify a white salt that is in a bottle with no label on it.

8.6 Water of crystallisation

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A **hydrated salt** is a salt which contains water as part of its crystalline structure. Examples are hydrated copper(II) sulfate, $CuSO_4.5H_2O$, and hydrated cobalt(II) chloride, $CoCl_2.6H_2O$.

The water present in crystals of a hydrated salt is known as **water of crystallisation**.

A salt that does not contain water of crystallisation is called an **anhydrous salt**.

When hydrated salts are heated, the water is given off and an anhydrous salt is left behind.

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$

blue crystals — hydrated copper(II) sulfate → white powder — anhydrous copper(II) sulfate

$$CoCl_2.6H_2O(s) \rightarrow CoCl_2(s) + 6H_2O(g)$$

pink crystals — hydrated cobalt(II) chloride → blue powder — anhydrous cobalt(II) chloride

Sample questions

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1 For each of the soluble salts below:

i name two substances that can be used to prepare the salt [2]

ii state which method you would use. [1]

a zinc nitrate

b potassium chloride

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Student's answers

a i *aqueous zinc hydroxide and dilute nitric acid*
ii *titration*
b i *potassium carbonate and dilute hydrochloric acid*
ii *Add excess potassium carbonate to dilute hydrochloric acid.*

Teacher's comments

The student chose the correct acids.
a Zinc hydroxide is insoluble in water and therefore cannot be titrated with dilute nitric acid.
b The student chose the correct substances. However, potassium carbonate is one of the few soluble carbonates, so the chosen method cannot be used.

Correct answers

a i solid zinc hydroxide, zinc carbonate, zinc oxide or metallic zinc, and dilute nitric acid
ii Add excess solid to the dilute nitric acid.
b i dilute hydrochloric acid and potassium carbonate or potassium hydroxide
ii titration

2 a Write word equations for the reactions between dilute hydrochloric acid and
i zinc
ii magnesium carbonate
iii calcium oxide

b Write balanced equations for the reactions between sulfuric acid, H_2SO_4, and
i magnesium, Mg
ii copper(II) carbonate, $CuCO_3$
iii potassium hydroxide, KOH

Student's answers

a i *hydrochloric acid + zinc → zinc chloride + water*
ii *hydrochloric acid + magnesium carbonate → magnesium chloride + water + carbon dioxide*
iii *hydrochloric acid + calcium oxide → calcium chloride + hydrogen*

b i $H_2SO_4 + 2Mg \rightarrow Mg_2SO_4 + H_2$
ii $H_2SO_4 + CuCO_3 \rightarrow CuSO_4 + H_2O + CO_2$
iii $H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$

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Teacher's comments

a i acid + metal → salt + hydrogen, not water
ii The student's answer is correct.
iii acid + base → salt + water, not hydrogen

b i Magnesium sulfate is $MgSO_4$ not Mg_2SO_4.
This is because the charge on a magnesium ion is 2+ and not 1+.
ii The student's answer is correct.
iii The equation is not balanced.

Correct answers

a i hydrochloric acid + zinc → zinc chloride + hydrogen
ii hydrochloric acid + magnesium carbonate → magnesium chloride + water + carbon dioxide
iii hydrochloric acid + calcium oxide → calcium chloride + water

b i $H_2SO_4 + Mg \rightarrow MgSO_4 + H_2$
ii $H_2SO_4 + CuCO_3 \rightarrow CuSO_4 + H_2O + CO_2$
iii $H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$

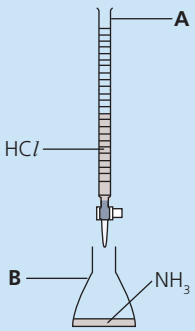
Exam-style questions

1 There are two general methods for the preparation of soluble salts.
Method 1: Adding an excess of an insoluble base or insoluble carbonate or metal to a dilute acid.
Method 2: Titration using an acid and an alkali or a soluble carbonate.
For each of the following salt preparations:
i Choose Method 1 or Method 2.
ii Name any additional reagent which is required.
iii Write the equation.

a cobalt(II) chloride starting with the insoluble compound cobalt(II) carbonate [4]
b potassium nitrate from aqueous potassium hydroxide [3]
[Total: 7]

2 Give full experimental details of how you would make pure dry crystals of magnesium sulfate starting with magnesium carbonate. You should include an equation in your answer. [Total: 10]

3 A student titrated hydrochloric acid with aqueous ammonia using the apparatus below.



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a Name apparatus **A** and apparatus **B**. [2]

b A titration is carried out in order to measure the exact volume of hydrochloric acid that is required to neutralise the ammonia in apparatus **B**.

i Name the type of substance that should be added to the ammonia before the acid is added. [1]

ii Give an example of the type of substance you have given in **(b)(i)**. [1]

c Describe how the pH of the solution in **B** changes as hydrochloric acid is added to the flask. [2]

d Complete the word and symbol equations for this reaction.

ammonia + hydrochloric acid → _____

_____ + $HCl \rightarrow NH_4Cl$ [2]

[Total: 8]

4 You are provided with a mixture of scandium oxide and copper(II) oxide.
Scandium oxide is an amphoteric oxide and copper(II) oxide is a basic oxide. Describe how you could obtain a sample of pure copper(II) oxide from the mixture. Both solids are insoluble in water. [Total: 5]

5 Lead(II) iodide is made by precipitation because it is insoluble in water.
You are provided with solid lead(II) nitrate and solid sodium iodide.
Describe how you would make a *pure* sample of lead(II) iodide by precipitation.
Your answer should include:
- practical details
- an ionic equation, with state symbols, for the reaction [Total: 9]

6 If you have not revised carboxylic acids, see Section 13.3.
Ethanoic acid is a weak acid. It is also an organic acid.
Ethanoic acid can be obtained from ethanol. A dilute solution of ethanoic acid is commercially available as vinegar.

a i State the meaning of the term *weak* with reference to acids. [1]

ii State the meaning of the term *acid* with reference to protons. [1]

iii Write an equation, with state symbols, for the dissociation of ethanoic acid showing that it is a weak acid. [3]

b i Name the type of substance that reacts with ethanol in order to convert it into ethanoic acid. [1]

ii Name the substance that reacts with ethanol when it is converted into vinegar. [1]

iii Write a chemical equation for the reaction that occurs in **(b)(ii)**. State symbols are not required. [1]

[Total: 8]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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9 The Periodic Table

Key objectives

By the end of this section, you should be able to:

- describe the Periodic Table as an arrangement of elements in periods and groups and in order of increasing proton number/atomic number
- describe the change from metallic to non-metallic character across a period
- describe the relationship between group number and the charge of the ions formed from elements in that group
- explain similarities in the chemical properties of elements in the same group
- explain how the position of an element in the Periodic Table can be used to predict its properties
- know that:
  - the number of outer shell electrons in an atom is equal to the group number in Groups I to VII
  - the number of occupied electron shells in an atom is equal to the period number
  - Group VIII atoms (noble gases) have a full outer shell of electrons
- describe the Group I alkali metals, lithium, sodium and potassium, as relatively soft metals
- describe how melting point, density and reactivity change down Group I
- predict the properties of other elements in Group I
- describe the Group VII halogens, chlorine, bromine and iodine, as diatomic non-metals and their appearance at r.t.p
- describe how density and reactivity change down Group VII
- describe and explain the displacement reactions of halogens with other halide ions
- predict the properties of other elements in Group VII
- describe the Group VIII noble gases as monatomic gases and explain this and their reactivity in terms of electronic configuration
- describe the transition elements as metals and know their general properties (densities, melting points, colour of compounds, catalytic behaviour)

- identify trends in groups, given information about the elements

- understand that transition metal ions have variable oxidation numbers

Key terms

REVISED

Term	Definition
Alkali metals	The six metallic elements in Group I of the Periodic Table.
Electronic configuration	A shorthand method of describing the arrangement of electrons within the electron shells of an atom.
Group	A vertical column of elements in the Periodic Table containing elements with the same number of electrons in their outer shell.
Halogens	The elements found in Group VII of the Periodic Table.
Noble gases	The elements found in Group VIII of the Periodic Table.
Periodic Table	A table of elements arranged in order of increasing proton number.
Periods	The horizontal rows of elements in the Periodic Table. The atoms of elements in a period have the same number of occupied shells.
Transition elements	The elements found in the centre of the Periodic Table, between Groups II and III.

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9.1 Development of the Periodic Table

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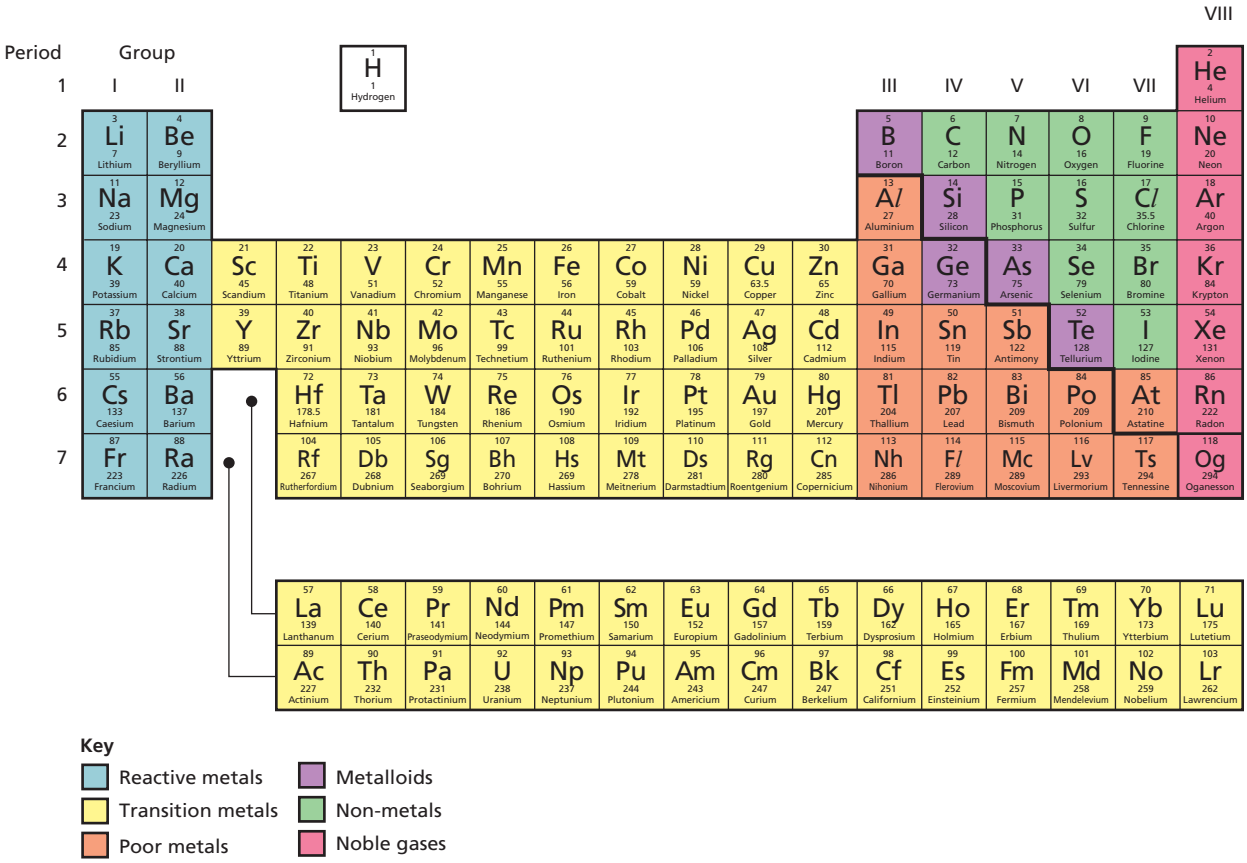
Early tables of elements arranged them in order of mass number or relative atomic mass. However, the **Periodic Table** contains the elements arranged in order of increasing proton number/atomic number.

(Relative atomic mass does tend to increase as proton number increases, but there are places where it decreases, for example argon to potassium.)

- The vertical columns of elements are called **groups**.
- The horizontal rows of elements are called **periods**.

Across Periods 2 and 3, there is a gradual change from metals on the left-hand side to non-metals on the right-hand side.



▲ **Figure 9.1 The modern Periodic Table**

9.2 Electronic configuration and the Periodic Table

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Atoms of elements in the same group have the same number of electrons in the outer shell. The number of electrons in the outer shell determines the chemical properties of the element.

The number of occupied shells present in an atom of an element is the same as the period number.

For example, potassium has proton number 19, therefore its **electronic configuration** is 2,8,8,1.

- There is one electron in the outer shell, which means potassium is in Group I of the Periodic Table.
- Potassium has electrons in four occupied shells, which means it is in Period 4.

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▼ **Table 9.1 Group number and electronic configuration**

Group number	Number of outer shell electrons	Change to electronic configuration when ions form	Charge on the ions
I	1	Lose 1 electron	1+
II	2	Lose 2 electrons	2+
III	3	Lose 3 electrons	3+
IV	4		
V	5	Gain 3 electrons	3–
VI	6	Gain 2 electrons	2–
VII	7	Gain 1 electron	1–

Skills

Where does an element belong?

Worked example

Sulfur has a proton number of 16.

State in which group and period of the Periodic Table sulfur is found.

Explain how you deduced your answers.

Answer

All atoms contain equal numbers of protons and electrons.

Therefore, a sulfur atom contains 16 electrons.

16 electrons give an electronic configuration of 2,8,6.

The group number is the same as the number of electrons in the outer shell.

Therefore, sulfur is in Group VI (6).

The period number is the number of shells that contain electrons.

Therefore, sulfur is in Period 3.

9.3 Group I – the alkali metals

REVISED

The Group I elements are known as the **alkali metals** because they react with water to produce alkaline solutions. The Group I elements are *very reactive* metals.

In order of increasing proton number, the Group I elements are lithium, sodium, potassium, rubidium, caesium and francium. Only lithium, sodium and potassium are found in school laboratories because rubidium, caesium and francium are dangerously reactive. Francium is also radioactive.

Properties

Group I elements:

- are stored under oil because they react rapidly with oxygen in the air
- are good conductors of heat and electricity
- can be cut with a knife because they are soft
- are shiny when cut, but tarnish rapidly due to reaction with oxygen in the air
- have low densities, melting points and boiling points compared to transition metals

Densities increase as you move down Group I.

The melting points and boiling points also increase down the group.

Reaction with water

All Group I elements react *vigorously* with water at room temperature.

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The reactivity of the Group I metals increases down the group. If rubidium and caesium are added to water, an explosive reaction occurs, which is why they are not kept in school laboratories.

Exam questions often ask for observations or ask *What would you see ... ?*

Observations you might make in this case are:

- the metal disappears
- sodium disappears more rapidly than lithium
- the metal melts
- bubbles/fizzing/effervescence (these all effectively mean the same thing)
- the metal floats and moves around on the surface of the water
- sodium moves around the surface faster than lithium
- potassium bursts into a lilac flame

However, the following are *not* observations:

- names of the products
- a gas is given off (it is not possible to see a colourless gas)
- an alkaline solution forms (it is not possible to see that a solution is alkaline by observation alone)
- colour change of an indicator (unless an indicator is mentioned in the question)

The equation for the reaction of sodium with water is:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

The equations with the other Group I metals are the same (including balancing numbers) – just replace Na with the symbol for the other metals.

9.4 Group VII – the halogens

REVISED

The Group VII elements are known as the **halogens**.

In order of increasing proton number, they are fluorine, chlorine, bromine, iodine and astatine. Only chlorine, bromine and iodine are found in school laboratories (see below for the reasons for this).

The Group VII elements are all non-metallic and exist as diatomic molecules (molecules containing two atoms). The appearances of those found in schools are shown in Table 9.2.

▼ **Table 9.2 Physical appearance of chlorine, bromine and iodine**

Element	Appearance at r.t.p.
Chlorine	Pale yellow–green gas
Bromine	Red–brown liquid
Iodine	Grey–black solid

The colours become *darker* as you move down the group. The change in physical state from gas to liquid to solid down the group indicates an increase in density down the group (due to an increase in the strength of intermolecular forces).

This means we can use Table 9.2 to predict the properties of astatine and fluorine. For example, fluorine will be a gas at room temperature and pressure (r.t.p.) and astatine will be a solid at r.t.p.

Revision activity

Create a table to compare and contrast how the properties of elements from Groups I and VII change as you move down each group. Consider melting point, boiling point, reactivity and anything else you think is important.

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Halogen displacement reactions

Chlorine displaces bromine from an aqueous solution of potassium bromide, turning the colourless solution to orange–yellow. The equation for this is:

$$Cl_2(g) + 2KBr(aq) \rightarrow Br_2(aq) + 2KCl(aq)$$

Table 9.3 shows the observations made, and the reasons for them, when halogens (or solutions of a halogen in water) are added to colourless aqueous solutions of potassium halides (chlorides, bromides and iodides).

▼ **Table 9.3 Halogen displacement reactions**

	Aqueous potassium chloride, KCl	Aqueous potassium bromide, KBr	Aqueous potassium iodide, KI
Chlorine, Cl_2		Solution turns orange–yellow (bromine produced)	Solution turns brown (iodine produced)
Bromine, Br_2	No change		Solution turns brown (iodine produced)
Iodine, I_2	No change	No change	

As can be seen from the table:

- Chlorine displaces bromine and iodine.
- Bromine displaces iodine, but does not displace chlorine.
- Iodine does not displace chlorine or bromine.

Halogens higher up the group can displace those lower down, indicating that the reactivity of the halogens decreases down the group.

Alternatively, we can say that reactivity increases up Group VII.

This is opposite to the trend in reactivity shown in Group I.

We can use this information to make predictions about other halogens and halides (see the questions at the end of the chapter).

However:

- Reactions involving fluorine only occur in theory because, in practice, fluorine reacts violently with water so cannot be used.
- Astatine is radioactive and cannot be used.

9.5 Group VIII – the noble gases

REVISED

The Group VIII elements are known as the **noble gases**. In order of increasing proton number, they are helium, neon, argon, krypton, xenon and radon.

The Group VIII elements are all:

- colourless gases
- monatomic – their atoms all have a full outer shell of electrons, so do not form covalent bonds creating diatomic molecules
- *very unreactive* because they have a full outer shell of electrons without sharing, losing or gaining electrons in a chemical reaction

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9.6 Transition elements

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Transition elements are all metals, so they are sometimes referred to as **transition metals**. They are found in the wide section of the Periodic Table between Groups II and III. Common examples are copper, iron and nickel.

Physical properties

Transition elements have the physical properties of 'typical' metals (see Chapter 10). In addition, transition elements have:

- high melting points
- high densities

Chemical properties

- Transition elements form coloured compounds. For example, copper(II) sulfate crystals are blue and potassium manganate(VII) is purple.
- The elements and their compounds show catalytic activity. For example, iron is used in the Haber process and vanadium(V) oxide is used in the Contact process.

- Transition elements have variable oxidation states. For example, iron can form Fe^{2+} and Fe^{3+} ions.
  - In Fe^{2+}, iron has an oxidation number of +2.
  - In Fe^{3+}, iron has an oxidation number of +3.

Revision activity

There are many similar facts to learn in this chapter. If music helps you to concentrate, try playing the same song or tune every time you revise Group I, a different song for Group VII, and so on. (If you find background noise distracting, this isn't the method for you.)

9.7 The position of hydrogen

REVISED

Hydrogen is not placed in any of the groups of the Periodic Table.

A hydrogen atom has one electron in its outer shell. It can lose this electron and become a H^+ ion. Therefore, hydrogen is similar to Group I elements, which lose one electron to form ions with a single positive charge. However:

- The Group I elements are solid and metallic hydrogen is gaseous and non-metallic.
- The Group I elements react vigorously with water, whereas hydrogen is insoluble in water and does not react.

As hydrogen atoms have one electron in the first shell, they only need to gain one electron to achieve a full outer shell. This is also true of Group VII elements. Hydrogen also forms diatomic molecules like the Group VII elements. However:

- Hydrogen is not coloured, whereas the Group VII elements have a variety of colours.
- Hydrogen does not take part in many of the reactions of Group VII elements.

Sample question

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Group I elements have one electron in their outer shell. They lose one electron to achieve a full outer shell.

Group VII elements have seven electrons in their outer shell. They gain one electron to achieve a full outer shell.

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A hydrogen atom contains one electron. It can gain one electron or lose one electron to achieve a full outer shell.

a Write the symbol for the particle that forms when a hydrogen atom:
- **i** gains an electron [1]
- **ii** loses an electron. [1]

b Give two pieces of evidence that suggest hydrogen should be present in Group I of the Periodic Table rather than in Group VII. [2]

c Give one piece of evidence that suggests hydrogen should be present in Group VII of the Periodic Table rather than in Group I. [1]

Student's answers

a **i** H^+
ii H^-

b
- *forms ions with a single positive charge*
- *reacts vigorously*

c *exists as a gas*

Teacher's comments

a Electrons have a negative charge. When an atom, X, gains an electron, it becomes X^-. When an atom loses an electron, it becomes X^+. The student has the charges the wrong way round.

b *Forms ions with a single positive charge* is a correct answer. *Reacts vigorously* is a meaningless statement because there is no reference to what it is reacting with. Group I elements react vigorously with water. Hydrogen does not.

c The statement *exists as a gas* is meaningless. All substances can exist as solids, liquids or gases depending on the temperature. Furthermore, in Group VII, fluorine and chlorine exist as gases at r.t.p., bromine exists as a volatile liquid and iodine exists as a solid.

Correct answers

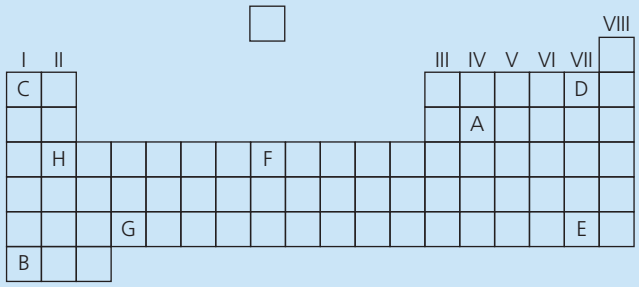
a **i** H^-
ii H^+

b
- forms ions with a single positive charge
- forms at the cathode in electrolysis

c exists as diatomic molecules (H_2)

Exam-style questions

1 The diagram below shows part of the Periodic Table.



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Use the letters **A** to **H** inclusive to answer the questions that follow. Each letter may be used once, more than once or not at all. Give the letter that represents:

a the Group I element that is most reactive [1]
b the Group VII element that is most reactive [1]
c a transition element [1]
d an element in Period 3 [1]
e an element whose atoms have four electrons in their outer shell [1]

[Total: 5]

2 Use the Periodic Table to predict reactions that would occur between:

a fluorine and aqueous potassium chloride
b astatine and aqueous potassium fluoride
c bromine and aqueous potassium astatide
d iodine and aqueous potassium fluoride

If you predict that a reaction would occur, write a chemical equation for the reaction. If you predict that a reaction would *not* occur, write no reaction. [Total: 6]

3 Vanadium is a transition element.
Vanadium is a good conductor of electricity.
Vanadium forms soluble salts.
Vanadium forms coloured compounds.
Vanadium(V) oxide is a catalyst.
Vanadium forms a basic oxide.
Vanadium has a very high density.

a Give two properties from the list above that show ways in which vanadium differs from Group I elements. [2]
b Give two properties from the list that show ways in which vanadium is similar to Group I elements. [2]

[Total: 4]

4 **a** Describe the trend in reactivity of Group I elements. [1]
b **i** State two observations that can be made when sodium is added to water. [2]
ii Write a chemical equation for the reaction that occurs when sodium is added to water. [2]

[Total: 5]

5 The Group VIII elements are called the noble gases.
Use your knowledge of electronic configuration, and your knowledge of ionic and covalent bonding, to explain why the noble gases do not show any chemical reactions. [3]

6 Copper and iron have variable oxidation states. State the formulae of:

a copper(I) oxide [1]
b copper(II) nitrate [1]
c iron(II) chloride [1]
d iron(III) sulfate [1]

[Total: 4]

7 Use the table of halogen displacement reactions (Table 9.3 on page 97) to help answer the following questions. Write chemical equations (with state symbols) and ionic equations for the reactions that occur between:

a chlorine and aqueous potassium iodide [4]
b bromine and aqueous potassium iodide [4]

[Total: 8]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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10 Metals

Key objectives

By the end of this section, you should be able to:

- compare the general physical properties of metals and non-metals, including:
  - thermal conductivity
  - electrical conductivity
  - malleability and ductility
  - melting points and boiling points
- describe the uses of metals in terms of their physical properties, including:
  - aluminium in aircraft, overhead cables and food containers
  - copper in electrical wiring
- describe the reactions of metals with:
  - dilute acids
  - cold water and steam
  - oxygen
- deduce an order of reactivity from a given set of experimental results
- state the order of the reactivity series as: potassium, sodium, calcium, magnesium, aluminium, carbon, zinc, iron, hydrogen, copper, silver, gold
- describe the reactions, if any, of:
  - potassium, sodium and calcium with cold water
  - magnesium with steam
  - magnesium, zinc, iron, copper, silver and gold with dilute hydrochloric acid

  and explain these reactions in terms of the position of the metal in the reactivity series

- describe the relative reactivities of metals in terms of:
  - their tendency to form positive ions
  - displacement reactions, if any, with the aqueous ions of magnesium, zinc, iron, copper and silver
- explain the apparent unreactivity of aluminium in terms of its oxide layer

- describe tests using aqueous sodium hydroxide and aqueous ammonia to identify the aqueous cations NH_4^+, Ca^{2+}, Cu^{2+}, Fe^{2+}, Zn^{2+}, Al^{3+}, Cr^{3+} and Fe^{3+}
- relate the ease by which metals are obtained from their ores to their position in the reactivity series
- describe the extraction of iron from hematite in the blast furnace

- state the symbol equations for the extraction of iron from hematite

- state the conditions required for the rusting of iron and steel
- state some common barrier methods to prevent rusting and describe how they work

- describe the use of zinc in galvanising
- explain sacrificial protection in terms of the reactivity series and electron loss

- describe alloys, including brass and stainless steel, as mixtures of a metal with other elements
- state that alloys are harder and stronger than pure metals
- describe the uses of alloys, including stainless steel in cutlery, in terms of their physical properties
- identify representations of alloys from diagrams

- explain in terms of structure why alloys can be more useful than pure metals

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Key terms

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Term	Definition
Alloy	A mixture of two or more metals or of a metal with a non-metal.
Corrosion	The process that takes place when metals and alloys are chemically attacked by oxygen, water or any other substance found in their immediate environment.
Metals	A class of chemical elements which have a characteristic lustrous appearance and are good conductors of heat and electricity.
Reactivity series of metals	An order of reactivity of metals, giving the most reactive metal first, based on results of the reactions of metals with oxygen, water and dilute hydrochloric acid.
Rust	An orange–brown layer of hydrated iron(III) oxide found on the surface of iron and steel.
Sacrificial protection	A method of rust prevention in which a layer of a more reactive metal is applied to the surface of iron or steel.

10.1 Properties of metals

REVISED

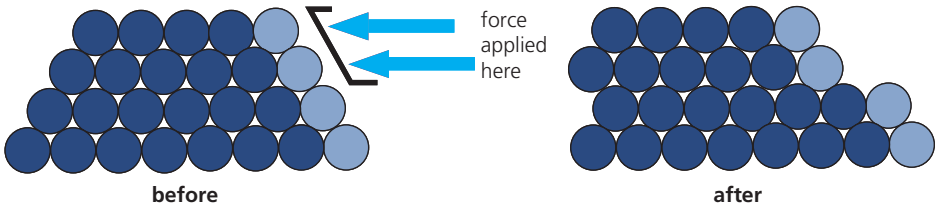
Physical properties

The physical properties of **metals** are shown in Table 10.1 (see also Chapter 9).

▼ **Table 10.1 Physical properties of metals and non-metals**

Physical property	Metal	Non-metal
Physical state at r.t.p.	Solid (except mercury)	Solid, liquid or gas
Malleability and ductility	Good	Poor
Melting point and boiling point	Usually high	Low for simple molecules High for giant covalent molecules
Conductivity (thermal and electrical) of solid	Good	Poor (except graphite)

Metals are malleable (can be hammered into different shapes) and ductile (can be drawn into wires). Although metallic bonds are strong, metals are not rigid because the ions are all the same size so the rows of ions can slide over each other when a force is applied.



▲ **Figure 10.1 The positions of the positive ions in a metal before and after a force has been applied**

10.2 Metal reactions

REVISED

Table 10.2 shows the differences in properties of the elements depending on their position in the **reactivity series**.

The elements are arranged with reactivity decreasing from the top to the bottom.

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▼ **Table 10.2 The reactivity series (carbon and hydrogen are not metals and are included only for comparison)**

Reactivity series	Reaction with dilute acid	Reaction with air/oxygen	Reaction with water	Ease of extraction	
Potassium (K) Sodium (Na)	Produce H_2 with decreasing vigour	Burn very brightly and vigorously	Produce H_2 with decreasing vigour with cold water	Difficult to extract	Increasing reactivity of metal ↑
Calcium (Ca) Magnesium (Mg)	↓	Burn to form an oxide with decreasing vigour	React with steam with decreasing vigour	Easier to extract	
Aluminium (Al)		↓	↓	↓	
[Carbon (C)]					
Zinc (Zn)					
Iron (Fe)					
[Hydrogen (H)]					
Copper (Cu)	Do not react with dilute acids	React slowly to form the oxide	Do not react with cold water or steam		
Silver (Ag)	↓	Do not react	↓		
Gold (Au)		↓		Found as the element (uncombined)	

Skills

Investigating metal reactions

With dilute acids

Metals *above* hydrogen in the reactivity series react with dilute hydrochloric acid to form a salt and hydrogen.

For example:

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$

The metal disappears and bubbles are seen. The reactions become less vigorous as we move down the reactivity series.

Copper, silver and gold do *not* react with water, steam or dilute acids as they are *below* hydrogen in the reactivity series.

It would be too dangerous to add potassium, sodium or calcium to a dilute acid as the reactions are far too vigorous and, therefore, extremely hazardous.

With water

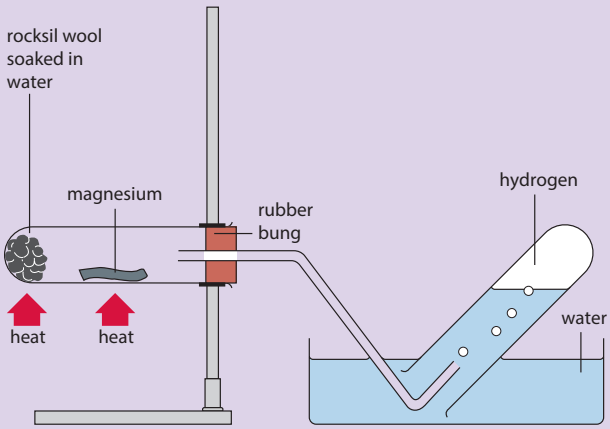
All Group I elements, including potassium and sodium, react vigorously with cold water at room temperature. The reactions are usually carried out in a glass trough. The observations are described in Section 9.3.

Calcium also reacts with cold water, but the reaction is not so vigorous as when potassium or sodium are used.

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

Magnesium reacts extremely slowly with cold and hot water. However, if steam is passed over heated magnesium, the magnesium glows very brightly and a white solid, magnesium oxide, remains in the test-tube. Hydrogen gas escapes and can be collected over water.

$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$



▲ **Figure 10.2 Apparatus used to investigate how metals such as magnesium react with steam**

With oxygen

All the metals listed, except for silver and gold, can be burned in oxygen. The metal oxide is the only product. For example:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

The reactions become less vigorous as you move down the reactivity series.

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Revision activity

Make a card for each of the elements in the reactivity series. Divide the cards between yourself and a friend. The player with 'potassium' lays the card down. The other player puts down the card which they think comes next in the series or passes if they do not have the right card. Continue until all the cards have been played in the right order.

10.3 Reactivity of metals and their uses

REVISED

▼ **Table 10.3 Uses of common metals**

Metal	Used to manufacture	Reason for use
Aluminium	Aircraft	Low density
	Overhead electrical cables	Low density Good electrical conductivity
	Food containers	Resistance to corrosion
Copper	Electrical wiring	Good electrical conductivity High ductility

Skills

Displacement reactions

The results of tests made by adding a metal to an aqueous solution containing ions of another metal, or by heating a metal with the oxide of another metal, can be used to put metals in order of reactivity.

More reactive metals will displace *less reactive* metals from aqueous solutions of their ions. For example, magnesium ribbon will displace copper from an aqueous solution of a salt, such as copper(II) sulfate solution.

In this example, the observations are:

- The magnesium ribbon disappears.
- The blue solution turns colourless.

magnesium + copper(II) sulfate solution → magnesium sulfate solution + copper

$$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$$

The ionic equation is:

$$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$$

The reaction occurs because magnesium has a greater tendency to form positive ions than copper.

If copper is added to a solution containing magnesium ions, such as aqueous magnesium sulfate, there is no reaction.

Unexpected behaviour of aluminium

Aluminium appears between magnesium and carbon in the reactivity series. However, aluminium often appears to be much less reactive than its position in the reactivity series suggests.

For example, if aluminium is placed in an aqueous solution of copper(II) sulfate, there is hardly any reaction.

This is because aluminium is so reactive that it reacts with the oxygen in the air, forming a layer of aluminium oxide which adheres to the aluminium underneath and protects the metal. Such a layer can be deliberately placed onto the surface of aluminium metal by a process called **anodising**. This means that aluminium can be used for things which would not normally be associated with reactive metals, such as aeroplane bodies, cooking foil, and pots and pans.

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10.4 Identifying metal ions

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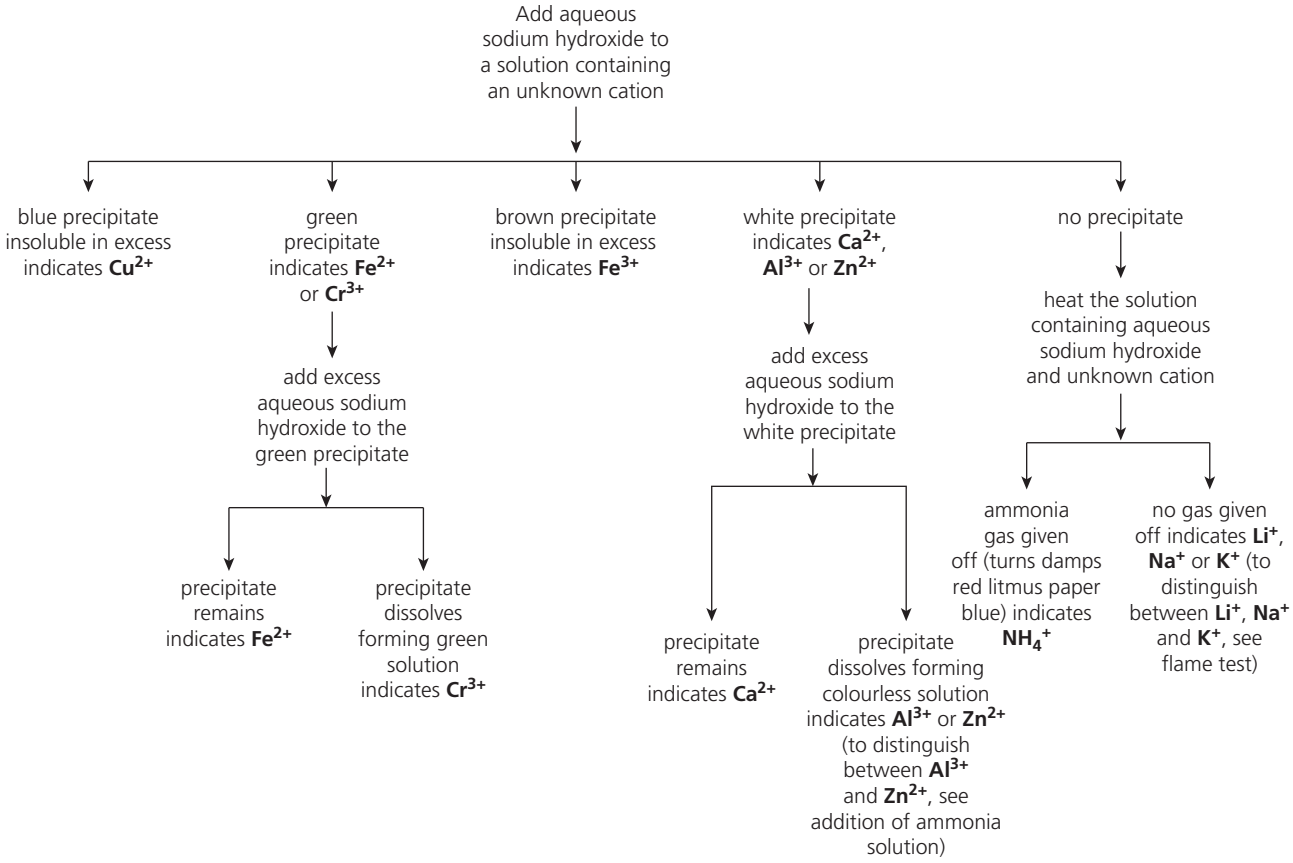
Cations (positive ions) can be identified using aqueous sodium hydroxide as shown in Figure 10.3. Other methods of identifying cations are described in Section 14.3.



▲ **Figure 10.3 Testing for cations (positive ions) in aqueous solution using aqueous sodium hydroxide**

Revision activity

Extend the poster or infographic you created for the revision activity in Section 8.5 (page 89) using the information in Figure 10.3.

10.5 Extraction of metals

REVISED

Metals can be extracted from their ores more easily as you go down the reactivity series.

There are three general methods of extracting metals from their ores:

1. Metals of low reactivity, for example copper, are extracted by:
   - **chemical reduction** using carbon/carbon monoxide as reducing agents or
   - **electrolysis** of aqueous solutions containing their ions
2. Metals of average reactivity, for example iron and zinc, are extracted by **chemical reduction** using carbon/carbon monoxide as reducing agents.
3. Very reactive metals, for example potassium, sodium, calcium, magnesium and aluminium, cannot be extracted by:
   - reduction because the ores are not reduced by chemical reducing agents such as carbon, carbon monoxide or hydrogen
   - electrolysis of aqueous solutions because hydrogen is formed at the cathode instead of the metal (see Chapter 5)

   Therefore, these metals are extracted by electrolysis of molten ionic compounds.

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Extraction of iron

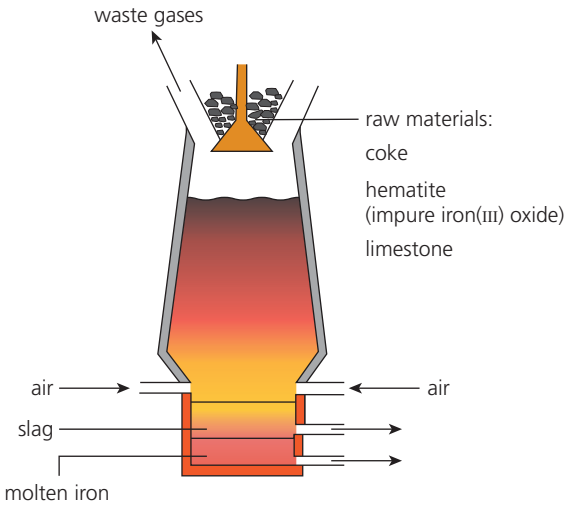
Iron is extracted from hematite (impure iron(III) oxide, Fe_2O_3) in a blast furnace.

- Hematite, coke, C, and limestone, $CaCO_3$, are fed into the top of the blast furnace.
- A blast of hot air enters near the bottom of the furnace.
- The coke reacts with the oxygen in the air, forming carbon dioxide. The reaction is *highly exothermic* and provides the high temperature required for the other reactions.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

- The carbon dioxide reacts with more coke higher up to produce carbon monoxide in an endothermic reaction.

$$C(s) + CO_2(g) \rightarrow 2CO(g)$$



▲ **Figure 10.4 A blast furnace**

- The carbon monoxide reduces the iron(III) oxide to molten iron.

$$Fe_2O_3 + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

- The molten iron trickles down to the bottom of the furnace and is tapped off.
- The function of the limestone is to remove the main impurity in the iron ore, which is silicon dioxide (silicon(IV) oxide).
  - The limestone thermally decomposes at the high temperature inside the blast furnace.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

  - Calcium oxide then reacts with silicon(IV) oxide to form calcium silicate, which forms a molten slag as a separate layer above the molten iron (it is less dense than iron).

$$CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$$

Slag is used by builders and road makers for foundations.

The iron produced in the blast furnace is called *pig iron* or *cast iron*. It contains about 4% carbon and its use is limited because it is brittle. The majority of pig iron is converted into steel.

Extraction of aluminium from bauxite

Aluminium is extracted from bauxite (impure aluminium oxide, Al_2O_3), as described in Section 5.3.

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10.6 Metal corrosion

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Corrosion is the process that takes place when metals or **alloys** react with oxygen, water or any other substance in their immediate environment. The metal or alloy is chemically changed and, therefore, its physical properties also change, making it less useful.

Rusting is a specific type of corrosion. Iron is the only metal that can form rust.

Rusting of iron

Rust can be described as hydrated iron(III) oxide, with a formula that can be represented as $Fe_2O_3.xH_2O$ (x is used because the amount of water of crystallisation varies from one sample of rust to another).

Iron only forms rust when it is exposed to oxygen (for example, from the air) and water.

Prevention of rusting

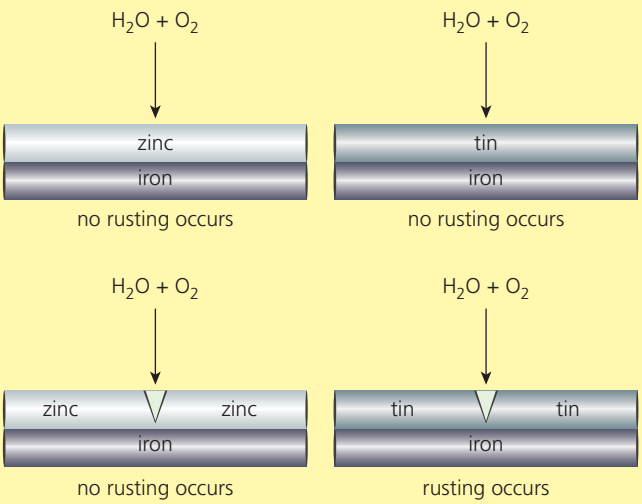
Rusting can be prevented by covering the iron with:

- paint
- oil or grease
- plastic
- a less reactive metal, such as tin – although this will only protect the iron if it is not scratched (see below)

These **barrier methods** prevent oxygen and water from coming into contact with the iron and stop a reaction from taking place.

Sacrificial protection

Some metals will continue to prevent iron from rusting even if the surface is scratched. Such metals must be above iron in the reactivity series but must not be so reactive that they will react rapidly with water themselves. When zinc is used for this purpose, the process is known as **galvanising** (Figure 10.5). Magnesium may also be used for **sacrificial protection** in this way.



▲ **Figure 10.5 Sacrificial protection**

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The first stage of rusting is the oxidation of iron to iron(II) ions by oxygen in the presence of water:

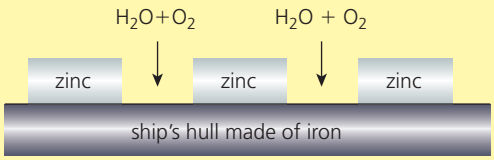
$$Fe \rightarrow Fe^{2+} + 2e^-$$

Zinc is a more reactive metal than iron – it forms positive ions by loss of electrons more readily than iron does. So, if zinc is present, the zinc will be oxidised in preference to the iron:

$$Zn \rightarrow Zn^{2+} + 2e^-$$

The electrons travel from the zinc to the iron. The iron does not lose electrons, which means that the iron is not oxidised – the first stage of rusting does not occur.

This will happen even if the zinc is scratched or does not completely cover the iron. Therefore, bars of zinc attached to the hull of a ship are sufficient to prevent it from rusting.



▲ **Figure 10.6 Sacrificial protection for the hull of a ship**

A common mistake when answering an exam question is to say that the zinc rusts instead of the iron or steel. Iron is the only metal that can form rust.

If tin is used instead of zinc and there is a scratch, the more reactive iron will be oxidised in preference to the tin. Thus, tin and other metals *below* iron in the reactivity series only prevent rusting when they are *not* scratched.

10.7 Alloys

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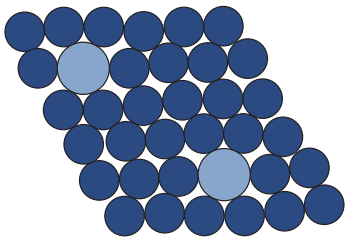
When a metallic object is required to be particularly strong, an **alloy** is often used instead of a pure metal. In alloys, such as brass, bronze and steel, the metallic element is mixed with small amounts of another element or elements.

Alloys are harder and stronger than pure metals, so are more useful.

▼ **Table 10.4 Uses of common alloys**

Alloy	Components	Use	Reason for use
Brass	Copper and zinc	Musical instruments	Hard Malleable
Stainless steel	Iron and other elements, such as chromium, nickel and carbon	Cutlery	Hard Resistant to corrosion



▲ **Figure 10.7 Alloy structure**

Figure 10.7 shows that the ions or atoms of the other elements in an alloy are a different size to those of the main element.

This size difference prevents the layers of metallic ions from *sliding* over each other and results in increased strength and hardness. In a metallic element, the particles are all the same size, which means the layers *can* slide over each other. Therefore, an alloy retains its shape much better than a pure metal when a force is applied.

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Sample questions

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1 You are provided with a mixture of powdered copper and powdered zinc. Describe how you would obtain a sample of pure copper from the mixture. You should give all observations for any reactions that you describe. [4]
Note: neither metal dissolves in water.

Student's answer

- *Add dilute hydrochloric acid to the mixture.*
- *Filter off the copper.*
- *Wash the copper and dry in a low oven.*

Teacher's comments

- An *excess* of dilute hydrochloric acid must be used in order to ensure that all the zinc reacts. The student should explain that the bubbling stops when all the zinc has reacted.
- Dilute sulfuric acid could be used instead. Dilute nitric acid should be avoided as some copper may react as well as the zinc.
- The mixture should be stirred and heated to increase the rate of reaction.
- After filtration, the copper should be washed with distilled water before it is dried.

Correct answer

- Add an excess of dilute hydrochloric acid to the mixture.
- Stir and warm the mixture. The zinc reacts and dissolves, creating bubbles of hydrogen.
- When no more bubbles are seen, filter off the copper.
- Wash the copper with distilled water and dry in a low oven.

2 Write the following chemical equation as an ionic equation:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Student's answer

$Zn^{2+}(s) + Cu(aq) + SO_4^{2-}(aq) \rightarrow Zn(aq) + SO_4^{2-}(aq) + Cu^{2+}(s)$

Teacher's comments

The element zinc contains zinc atoms, Zn, and not Zn^{2+}. When zinc reacts with copper(II) sulfate, Zn changes into Zn^{2+}.

The copper ions in copper(II) sulfate, Cu^{2+}, change into Cu when copper(II) sulfate reacts with zinc.

The $SO_4^{2-}(aq)$ ions are spectator ions – they are unchanged and should not be present in the ionic equation.

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Note that:

- *All* ionic equations for displacement reactions between metals and metal ions where X is a more reactive metal than Y are of the type:

$$X(s) + Y^{2+}(aq) \rightarrow X^{2+}(aq) + Y(s)$$

- If the charges on the ions of the two metals are not the same, balancing needs to be carried out.

Correct answer

$SO_4^{2-}(aq)$ is the same on both sides and is cancelled out:

$$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

Final ionic equation:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Exam-style questions

1 Construct word equations for the reactions in which:

a dilute nitric acid reacts with magnesium [1]

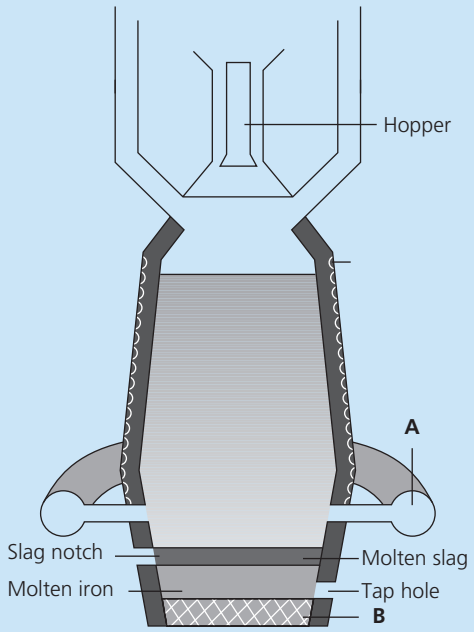
b chlorine reacts with aqueous potassium bromide [1]

c iron(III) oxide is reduced by carbon monoxide in a blast furnace [1]

d silicon dioxide is converted into calcium silicate in a blast furnace. [1]

[Total: 4]

2 Iron is extracted from its main ore in a blast furnace.



a i Name the main ore of iron used in the blast furnace. [1]

ii Name the substance that enters the blast furnace at **A**. [1]

iii Give two reasons for using coke in the blast furnace. [2]

b i Name the two products formed when the limestone decomposes. [2]

ii Name the substance that leaves the blast furnace at **B**. [1]

[Total: 7]

3 Steel can be protected from rusting by coating the steel with another metal, such as zinc.

a Name the element in steel that forms rust. [1]

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b Name the substances that react with the element in **(a)** to form rust. [1]

c Name two other substances that can be used instead of a metal to protect steel from rusting. [2]

d State how the substances named in **(c)** protect the steel from rusting. [1]

[Total: 5]

4 The results of some experiments carried out by adding a metal to aqueous solutions containing ions of another metal are shown in the table below, where ✓ means a reaction occurs and ✗ means no reaction occurs.

	$A(NO_3)_2(aq)$	$B(NO_3)_2(aq)$	$C(NO_3)_2(aq)$	$D(NO_3)_2(aq)$
Metal A(s)		✗	✗	✓
Metal B(s)	✓		✓	✓
Metal C(s)	✓	✗		✓
Metal D(s)	✗	✗	✗	

a Put the four metals in order of reactivity, starting with the most reactive first. [1]

b Write a chemical equation for the reaction occurring when metal B is added to $A(NO_3)_2(aq)$. [1]

c Write an ionic equation for the reaction occurring when metal C is added to $D(NO_3)_2(aq)$. [1]

[Total: 3]

5 Lithium is added to cold water in a glass trough.

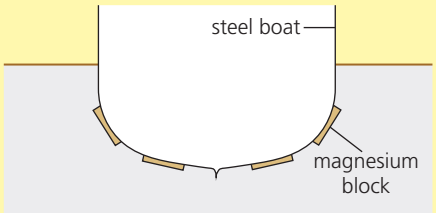
a Give three observations you would expect to make. [3]

b Write a chemical equation for the reaction that occurs. Include state symbols. [3]

c Methyl orange is added to the liquid in the trough after the reaction.
Give the colour it would become. [1]

[Total: 7]

6 Magnesium blocks can be attached to the bottom of steel boats. The magnesium does not completely cover the steel.



a State why the steel that is covered by magnesium does not form rust. [2]

b **i** Explain in terms of the formation of positive ions and loss of electrons why the magnesium prevents steel from rusting in the regions that are not covered. [2]

ii State the name given to the type of protection described in **(b)(i)**. [1]

c Name the method of protection if zinc is used instead of magnesium. [1]

d Explain why replacing the magnesium blocks with copper blocks will not prevent the boat from rusting where it is not covered. [1]

[Total: 7]

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11 Chemistry of the environment

Key objectives

By the end of this section, you should be able to:

Water

- describe chemical tests for the presence of water
- describe how to test for the purity of water
- explain why distilled water is used in practical chemistry rather than tap water
- state:
  - which substances may be present in water obtained from natural sources
  - the beneficial effects and harmful effects of these substances
- describe the treatment of water for the domestic water supply in terms of:
  - sedimentation and filtration
  - use of carbon
  - chlorination

Artificial fertilisers

- state that ammonium salts and nitrates are used as fertilisers
- describe the use of NPK fertilisers

Air and atmospheric pollution

- state the composition of clean, dry air
- state the sources of common pollutants in the air
- state the adverse effects of these common pollutants

- explain how oxides of nitrogen form in car engines

- describe photosynthesis
- state the word equation for photosynthesis

- state the symbol equation for photosynthesis
- describe how greenhouse gases, carbon dioxide and methane cause global warming

- state and explain strategies to reduce the effect of:
  - climate change
  - acid rain

Key terms

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Term	Definition
Catalytic converter	A device for converting pollutant exhaust gases from cars into less harmful emissions.
Fertiliser	A chemical substance added to soil to replace mineral salts to make plants grow more healthily.
Photosynthesis	The process by which green plants synthesise carbohydrates from carbon dioxide and water using light as the energy source and chlorophyll as the catalyst.
Pollution	The modification of the environment by human influence.

11.1 Water

REVISED

Anhydrous cobalt(II) chloride or anhydrous copper(II) sulfate can be used to test for the presence of water. The colour changes shown in Table 11.1 occur with water or anything containing it (including all aqueous solutions). Therefore, these methods are *not* used as a test for *pure* water.

▼ **Table 11.1 Tests for water**

	Original colour	Final colour
Anhydrous cobalt(II) chloride	Blue	Pink
Anhydrous copper(II) sulfate	White	Blue

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Purity of water

The **purity** of a water sample can be determined by measuring the boiling point. Pure substances boil and melt at specific temperatures, as opposed to a range of temperatures. For water at atmospheric pressure, the boiling point is 100°C and the melting point is 0°C.

Distilled water

Distilled water is used in practical chemistry rather than tap water. Tap water contains more impurities than distilled water and therefore may contain the ions that are being tested for in analysis (see Chapter 14).

Water from natural sources

Water obtained from natural sources contains various substances. Some of these substances have beneficial effects, whereas others have harmful effects.

▼ **Table 11.2 Beneficial and harmful effects of substances in water from natural sources**

Substance	Beneficial effect	Harmful effect
Dissolved oxygen	Essential for aquatic life	
Dissolved metal ions	Some metal ions are necessary for health, e.g. calcium ions are necessary for healthy growth of bones and teeth	Some metal ions, e.g. cadmium and mercury, are toxic
Plastics		Death of aquatic life
Microbes		Microbes present in sewage cause diseases
Dissolved nitrates and phosphates		Nitrates and phosphates from agricultural waste and detergents lead to removal of oxygen from water

Water treatment

Exact processes used to make water suitable for drinking vary from region to region. Common steps include:

- Sedimentation: this results in smaller, undissolved particles sinking to the bottom of a tank.
- Filtration: this involves passing impure water through screens to filter out floating debris.
- Addition of carbon to remove unwanted tastes.
- Chlorination: small amounts of chlorine gas are added to kill microbes.

Do not make the mistake of saying that chlorine is added to purify the water. Pure water contains water molecules and nothing else, so water containing small amounts of chlorine is *not* pure.

Revision activity

Make a set of cards for the different types of water treatment. Write the name of one of the processes on one side of a card and the reason for that process on the other. Use the cards in the same way as the key word cards you made in Section 2.4 (page 14).

11.2 Artificial fertilisers

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Fertilisers are substances that are added to soil to supply nutrients that are essential for the healthy growth of plants. NPK fertilisers contain nitrogen, phosphorus and potassium, which are the three main elements required.

- Very few plants can utilise nitrogen from the air, so fertilisers containing ammonium salts, such as ammonium sulfate or ammonium nitrate, supply the nitrogen.
- Ammonium phosphate and potassium chloride can be added to supply other essential elements.

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Skills

Nitrogen content of a fertiliser

Worked example

Find the percentage of nitrogen by mass in ammonium sulfate, $(NH_4)_2SO_4$.

Relative atomic masses, A_r: N =14, H = 1, S = 32, O = 16

Answer

Multiply out to remove the brackets:

$$(NH_4)_2SO_4 = N_2H_8SO_4$$

Find the relative formula mass, M_r:

$$N_2H_8SO_4 = (14 \times 2) + (1 \times 8) + 32 + (16 \times 4) = 132$$

The formula shows that the compound contains 2 moles of N atoms. Find the mass of these:

$$2N = 2 \times 14 = 28$$

Find the percentage of nitrogen:

$$(28 \div 132) \times 100 = 21.21\%$$

11.3 The air

REVISED

Air is a mixture and, as with all mixtures, its composition can vary.

The approximate composition of clean, dry air is:

- 78% nitrogen
- 21% oxygen
- 0.04% carbon dioxide
- 1% argon

Very small amounts of other noble gases are also present.

Notice that air does *not* contain hydrogen.

Photosynthesis

Photosynthesis occurs in green plants. Carbon dioxide in the atmosphere reacts with water in the presence of chlorophyll (in plant leaves), using energy from sunlight to form glucose and oxygen:

$$\text{carbon dioxide} + \text{water} \rightarrow \text{glucose} + \text{oxygen}$$

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

This process is important because:

- oxygen and glucose are essential for respiration – the process by which almost all living things obtain energy
- it removes carbon dioxide (a greenhouse gas) from the atmosphere

Global warming

Some of the energy from the Sun that reaches the Earth is reflected back into space. The rest is absorbed, heating up the Earth and its atmosphere. The covalent bonds in carbon dioxide and methane molecules absorb thermal energy. Some, but not all, of the thermal energy is re-emitted and travels into space. The thermal energy that does not escape causes an increase in the Earth's temperature. This results in **global warming**, which leads to **climate change**.

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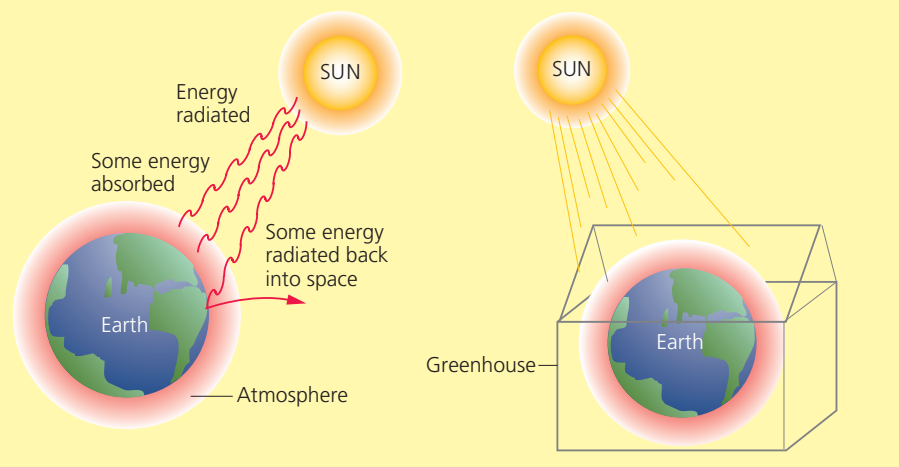

▲ **Figure 11.1 The greenhouse effect**

11.4 Atmospheric pollution

REVISED

Some common gaseous **pollutants**, their sources and the related adverse effects are shown in Table 11.3.

▼ **Table 11.3 Common atmospheric pollutants**

Pollutant	Source	Adverse effect
Carbon dioxide	Complete combustion of carbon-containing fuels, e.g. biomass and fossil fuels	Global warming, leading to climate change
Carbon monoxide	Incomplete combustion of carbon-containing fuels	Toxic
Particulates	Incomplete combustion of carbon-containing fuels	Respiratory problems and cancer
Methane	Decomposition of vegetation Waste gases from digestion in animals	Global warming, leading to climate change
Oxides of nitrogen	Car engines	Photochemical smog Respiratory problems Acid rain (caused by nitrogen dioxide)
Sulfur dioxide	Combustion of fossil fuels (particularly coal) containing sulfur compounds as impurities	Acid rain

Key points to note:

- Different pollutants cause different problems. Make sure you study Table 11.3 carefully and learn the sources of and problems caused by each individual pollutant.
- Sulfur dioxide does *not* come from the deliberate burning of sulfur – many fossil fuels contain small amounts of sulfur compounds as impurities.
- Oxides of nitrogen are produced by the reaction between nitrogen and oxygen, both of which come from the air, in car engines (not in the exhaust itself). The nitrogen is not present in the fuel.

Reducing the impact of atmospheric pollutants

Climate change

The amounts of carbon dioxide and methane (both greenhouse gases) in the atmosphere can be decreased by the strategies shown in Table 11.4.

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▼ **Table 11.4 Reducing carbon dioxide and methane in the atmosphere**

Strategy	Explanation
Planting more trees	Trees absorb carbon dioxide by photosynthesis
Reduction in livestock faming	Less methane is released by digestive processes in livestock
Decreasing use of fossil fuels	Less carbon dioxide is produced by complete combustion of carbon-containing fuels
Increasing use of alternative forms of energy (e.g. hydrogen, wind and solar)	Less carbon dioxide is produced by complete combustion of carbon-containing fuels

Acid rain

The amounts of sulfur dioxide and oxides of nitrogen in the atmosphere can be decreased by the strategies shown in Table 11.5.

▼ **Table 11.5 Reducing sulfur dioxide and oxides of nitrogen in the atmosphere**

Strategy	Explanation
Use of catalytic converters	Catalytic converters remove oxides of nitrogen
Use of low-sulfur fuels	Less sulfur dioxide is released into the atmosphere
Flue gas desulfurisation by calcium oxide	Less sulfur dioxide is released into the atmosphere because calcium oxide neutralises sulfur dioxide

Catalytic converters

Catalytic converters in vehicle exhausts remove pollutants that are produced by the engine which would otherwise enter the atmosphere.

One group of pollutants the catalytic converters remove is oxides of nitrogen, such as nitrogen monoxide, NO. If this gas is released into the atmosphere, it reacts with oxygen to produce nitrogen dioxide, NO_2. When nitrogen dioxide reacts with water and oxygen in the atmosphere, it forms a dilute solution of nitric acid, which is a constituent of acid rain.

Several reactions occur inside catalytic converters, including:

$$2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$$

This reaction removes both carbon monoxide and nitrogen monoxide.

The catalysts in catalytic converters include alloys containing transition elements, such as platinum, rhodium and palladium.

Revision activity

Create a mind map about atmospheric pollution. Make a branch for each gas and sub-branches showing where the gas comes from, the problems it causes and things that can be done to reduce this type of pollution and its effects.

Sample questions

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1 The following substances are all gases:

hydrogen **oxygen** **nitrogen** **carbon monoxide**
carbon dioxide **sulfur dioxide** **methane**

Use the names of the gases to match the descriptions below.
Each gas can be used once, more than once or not at all.

a burns in air to form water as the only product [1]
b is produced by the complete combustion of fossil fuels [1]
c is a hydrocarbon [1]
d is produced as a waste gas in the respiration of some animals [1]
e is removed from flue gas [1]
f makes up 78% of clean, dry air [1]

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Student's answers

a	methane	d	nitrogen
b	carbon dioxide	e	sulfur dioxide
c	methane	f	oxygen

Teacher's comments

a Methane does burn in air to produce water. However, carbon monoxide or carbon dioxide are also produced, so water is not the only product.
b Carbon dioxide is the correct answer.
c Methane is the correct answer.
d Nitrogen is exhaled when animals breathe out. However, the question asks for the waste gas that is *produced*, and nitrogen is not produced. It is already present in the air that is breathed in and passes unchanged through the body of animals before it is breathed out.
e Sulfur dioxide is the correct answer.
f The student has mixed up the percentages of oxygen and nitrogen, the two main gases in the air.

Correct answers

a	hydrogen	d	carbon dioxide
b	carbon dioxide	e	sulfur dioxide
c	methane	f	nitrogen

2 The word equation shows a reaction occurring in a catalytic converter.

nitrogen dioxide + carbon monoxide → nitrogen + carbon dioxide

a Choose one element from the following list which you think might be a suitable catalyst. Explain how you made your decision.
calcium **carbon** **copper** **sodium**
b State the adverse effect of oxides of nitrogen, such as nitrogen dioxide, NO_2, in the atmosphere.
c State the adverse effect of carbon monoxide in the atmosphere.
d Carbon dioxide is a product of the reaction occurring in a catalytic converter. State the adverse effect of the production of carbon dioxide.
e Write a symbol equation for the reaction occurring in a catalytic converter.

Student's answers

a Sodium – it is a reactive element.
b NO_2 is an atmospheric pollutant.
c Carbon monoxide is poisonous.
d CO_2 is an acidic gas and causes acid rain.
e $NO_2 + CO \rightarrow N + CO_2 + O$

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Teacher's comments

a Reactive elements are not usually used as catalysts. Catalysts are unchanged at the end of a reaction. Reactive elements are very unlikely to be unchanged.
b To state that NO_2 is an atmospheric pollutant is merely rewriting the question instead of answering it. A specific effect is required.
c Carbon monoxide could also be described as toxic.
d Carbon dioxide is not acidic enough to cause acid rain. The pH of acid rain is 4.0 or below, and carbon dioxide does not give rainwater such a low pH.
e The student has 'balanced' the equation by adding an extra oxygen atom on the right-hand side. You cannot balance an equation by changing formulae and/or adding symbols and/or formulae to either or both sides. You must write the correct formulae first and then use balancing numbers in front of the formulae.

Correct answers

a Copper is most likely to be a suitable catalyst because it is a transition element.
b Nitrogen dioxide causes acid rain. It is also a cause of photochemical smog and respiratory problems.
c Carbon monoxide is toxic or poisonous.
d Carbon dioxide is a greenhouse gas that causes global warming. Global warming leads to climate change.
e $2NO_2 + CO \rightarrow N_2 + CO_2$

Exam-style questions

1 **a** Write a word equation for the production of carbon dioxide in each of these reactions:
- **i** complete combustion of octane [1]
- **ii** thermal decomposition of calcium carbonate [1]
- **iii** reduction of iron(III) oxide by carbon monoxide [1]
- **iv** reaction between calcium carbonate and dilute hydrochloric acid. [1]

b Carbon dioxide can be removed from the Earth's atmosphere by photosynthesis.
- **i** Name the other reactant in photosynthesis. [1]
- **ii** Name the two products of photosynthesis. [2]
- **iii** State two conditions that are required for photosynthesis. [2]

[Total: 9]

2 Sulfur dioxide is an atmospheric pollutant.
- **a** State the source of sulfur dioxide in the atmosphere. [1]
- **b** State the adverse effect of sulfur dioxide in the atmosphere. [1]
- **c** Give three strategies by which the amount of sulfur dioxide is reduced. [3]

[Total: 5]

3 This question is about water treatment.
Draw lines on a copy of the diagram to **link the boxes** on the left with those on the right. The first one is done for you. [3]

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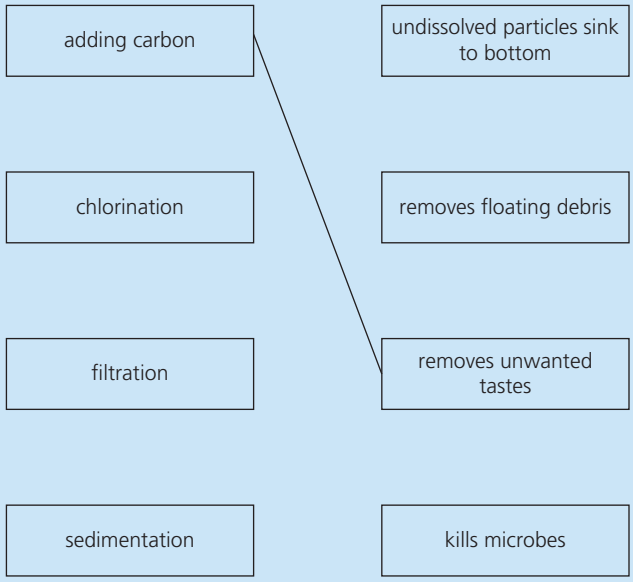

4 **a** Carbon dioxide is a product of the reactions below.
Write a symbol equation for each of these reactions. You may omit state symbols.

- **i** the complete combustion of decane, $C_{10}H_{22}$ [2]
- **ii** thermal decomposition of calcium carbonate [1]
- **iii** the fermentation of glucose (see Chapter 13) [2]
- **iv** the reaction between solid sodium carbonate and dilute hydrochloric acid [2]
- **v** the reaction between carbon monoxide and nitrogen monoxide in a catalytic converter [2]

b Carbon dioxide is a reactant in the reactions below.
Write a symbol equation for each of these reactions. You may omit state symbols.

- **i** the reaction between carbon dioxide and water in green plants [2]
- **ii** the reaction between carbon dioxide and coke in a blast furnace [2]

c Complete the following passage by adding the words below.

absorption **methane** **space** **thermal**

The greenhouse gases carbon dioxide and _____ cause global warming by the _____ of _____ energy, reducing the loss of thermal energy to _____ . [4]

[Total: 17]

5 Ammonium phosphate, $(NH_4)_3PO_4$, is used as a fertiliser.

a State the meaning of the term *fertiliser*. [1]

b Calculate the percentage of nitrogen, by mass, in ammonium phosphate, $(NH_4)_3PO_4$. [2]

[Total: 3]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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12 Organic chemistry 1

Organic chemistry is the study of covalent compounds containing carbon atoms bonded to atoms of hydrogen, oxygen, the halogens and nitrogen.

Key objectives

By the end of this section, you should be able to:

- state what is meant by the term functional group
- state what is meant by the term homologous series
- state what is meant by molecular formulae, displayed formulae and structural formulae
- draw the displayed formulae of methane, ethane, ethene and the products of their reactions referred to in this chapter

- describe and identify structural isomerism
- name and draw the structural and displayed formulae of unbranched alkanes and alkenes and the products of their reactions containing up to four carbon atoms per molecule as well as the structural isomers of C_4H_{10} and C_4H_8

- state the type of compound present given the chemical name ending in -ane, -ene or from a molecular, structural or displayed formula
- write and interpret the general formulae of alkanes and alkenes
- describe the bonding in alkanes and alkenes
- describe alkanes as being generally unreactive except in terms of combustion and substitution by chlorine

- describe what is meant by a substitution reaction
- describe the substitution reactions of alkanes with chlorine

- state the difference between saturated and unsaturated compounds in terms of carbon–carbon bonds
- state that alkanes are saturated hydrocarbons and alkenes are unsaturated hydrocarbons
- describe the tests for saturation and unsaturation

- describe what is meant by an addition reaction
- describe the chemical properties of alkenes in terms of addition reactions with:
  - bromine or aqueous bromine
  - hydrogen in the presence of a nickel catalyst
  - steam in the presence of an acid catalyst

- describe the manufacture of alkenes and hydrogen by cracking of larger alkane molecules
- describe the reasons for cracking larger alkane molecules
- define monomers and polymers

- identify the repeat unit in an addition polymer
- deduce the structure or repeat unit in an addition polymer from a given alkene and vice versa

- state that plastics are made from polymers
- describe how the properties of plastics have implications for their disposal
- describe the environmental challenges caused by plastics

Key terms

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Term	Definition
Addition reaction	A reaction in which an atom or group of atoms is added across a carbon–carbon double bond. In an addition reaction only one product is formed.
Alkane	A saturated hydrocarbon that contains single bonds only.
Alkene	An unsaturated hydrocarbon in which there is at least one carbon–carbon double bond.
Displayed formula	A formula showing all the atoms and bonds in one molecule of a compound.

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Term	Definition
Functional group	The atom or group of atoms responsible for the characteristic reactions of a compound.
Homologous series	A family of similar compounds with similar chemical properties and the same functional group and general formula that display a trend in physical properties. Each member differs from the rest by a –CH2– unit.
Hydrocarbon	A compound made of molecules containing carbon atoms and hydrogen atoms only.
Molecular formula	A formula showing the number of atoms of each element in one molecule of a substance.
Monomer	A simple molecule that can be polymerised.
Polymer	A substance possessing very large molecules consisting of repeated units.
Saturated hydrocarbon	A hydrocarbon in which the molecule has no double bonds.
Structural formula	A formula showing how groups of atoms are arranged in a molecule.
Structural isomerism	The existence of compounds with the same molecular formula but different structural formulae.
Substitution reaction	A reaction in which an atom or group of atoms is replaced by another atom or group of atoms.
Unsaturated hydrocarbon	A hydrocarbon in which there is at least one carbon–carbon double (or triple) bond.

12.1 Alkanes

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Hydrocarbons

Hydrocarbons are compounds made of molecules containing only carbon atoms and hydrogen atoms. The word *only* is important in this definition – compounds such as ethanol, C_2H_5OH, contain carbon and hydrogen but they are not hydrocarbons because they also contain oxygen.

Homologous series

A **functional group** is an atom or group of atoms responsible for the characteristic reactions of a compound.

Organic compounds belong to *families* of similar compounds known as **homologous series**, examples of which are alkanes, alkenes, alcohols, carboxylic acids and esters.

Members of a homologous series have:

- the same functional group
- the same general formula
- each member differing from the previous member by a $-CH_2-$ group of atoms
- similar chemical properties
- a trend in their physical properties, for example, melting points and boiling points that show almost constant increases between members of the series

Alkanes

Alkanes are members of a homologous series.

- The name of each member of the series ends in *-ane*.
- They have the general formula C_nH_{2n+2}.

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- They are **saturated hydrocarbons**, which means that all their bonds are single bonds (either C–C or C–H).
- They do not contain a functional group. The only bonds they contain (C–C and C–H) are found in all other organic compounds.

▼ **Table 12.1 First four unbranched members of the homologous series of alkanes**

Number of carbon atoms	Name	Molecular formula	Structural formula
1	Methane	CH_4	CH_4
2	Ethane	C_2H_6	CH_3CH_3
3	Propane	C_3H_8	$CH_3CH_2CH_3$
4	Butane	C_4H_{10}	$CH_3CH_2CH_2CH_3$

The names of alkanes are important because unbranched members of all other homologous series are named after the alkane with the same number of carbon atoms. Therefore, the names of all organic molecules with:

- one carbon atom begin with *meth-*
- two carbon atoms begin with *eth-*
- three carbon atoms begin with *prop-*
- four carbon atoms begin with *but-*

This does not apply to esters (see Chapter 13).

▼ **Table 12.2 Some organic compounds with two carbon atoms**

Alkane	Alkene	Alcohol	Carboxylic acid	Chloroalkane
Ethane	**Eth**ene	**Eth**anol	**Eth**anoic acid	Chloro**eth**ane

Revision activity

***M**y **e**lephant **p**lays **b**ongos* is a mnemonic for **m**eth-, **e**th-, **p**rop- and **b**ut-. Create one of your own – the sillier the better – to help you remember the beginnings of the names of organic molecules.

Formulae of organic compounds

Organic compounds have several different formulae.

- **Empirical formula**: This is the smallest whole number ratio of the atoms of each element in a compound (see Chapter 4).
- **Molecular formula**: This is the number of atoms of each element in one molecule of a substance (see Chapter 4). It gives no information about how the atoms are joined together.
- **Structural formula**: This shows how groups of atoms are arranged in a molecule.
- **Displayed formula**: This shows all the atoms and all the bonds in one molecule of a compound. Instead of asking for the displayed formula, exam questions sometimes ask you to draw the structure of a molecule showing all the atoms and all the bonds.

When you draw displayed formulae, make sure you have the right number of bonds (sticks) for each atom:

- All carbon atoms have four bonds.
- All hydrogen atoms have one bond.
- All oxygen atoms have two bonds.
- All halogen atoms have one bond.
- All nitrogen atoms have three bonds.

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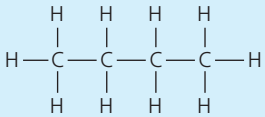
▼ **Table 12.3 Formulae of organic compounds, using butane as an example**

Compound	Empirical formula	Molecular formula	Displayed formula	Structural formula
Butane	C_2H_5	C_4H_{10}		$CH_3CH_2CH_2CH_3$

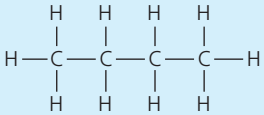
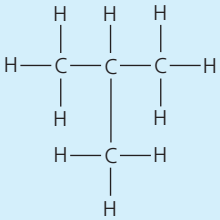
Structural isomerism

Structural isomerism is the existence of compounds with the *same* molecular formula but *different* structural formulae and, therefore, different displayed formulae.

It is easy to confuse the words *isotope* and *isomer*.

- Isotopes (see Chapter 2) are atoms of the same element with the same proton number but different nucleon numbers.
- Structural isomers are compounds with the same molecular formula but different structural formulae.

▼ **Table 12.4 Structural isomerism in butane**

Molecular formula	C_4H_{10}	C_4H_{10}
Displayed formula		
Structural formula	$CH_3CH_2CH_2CH_3$	CH_3CHCH_3 | CH_3 or $CH_3CH(CH_3)CH_3$
Name	Butane	2-methylpropane

The compound with the molecular formula C_4H_{10} has two structural isomers with different structural and displayed formulae (Table 12.4). As they are different compounds, they have different names.

- Butane is often referred to as a straight-chain or an unbranched molecule because the carbon atoms are arranged one after another.
- 2-methylpropane is often referred to as a branched-chain molecule.

The 2 shows which atom in the main chain the $-CH_3$ group (the methyl group) is joined to. In this case, the $-CH_3$ group can only be in position 2 (otherwise it just extends the main chain). Numbers are only essential when there are alternatives, e.g. 2-methylpentane and 3-methylpentane.

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12.2 The chemical behaviour of alkanes

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Alkanes are relatively unreactive because the single carbon–carbon bonds need a lot of energy to break.

Combustion

Alkanes undergo **combustion** in air or oxygen, producing energy, which is why alkanes are used as fuels.

Complete combustion occurs in excess oxygen. The products are carbon dioxide and water. For example:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Incomplete combustion of alkanes in a limited supply of air or oxygen leads to the production of (toxic) carbon monoxide as well as water (see Chapter 11):

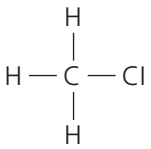
$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$$

Reaction with chlorine

It is not possible to add atoms to alkane molecules without first removing atoms. This type of reaction is called a **substitution reaction** because one atom or group of atoms is replaced by another atom or group of atoms.

When methane is reacted with chlorine in the presence of ultraviolet light, one chlorine atom replaces one hydrogen atom. The organic product is chloromethane, CH_3Cl:

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$$



▲ **Figure 12.1 Chloromethane**

The displayed formula of chloromethane is shown in Figure 12.1.

Unless the chlorine supply is limited, the reaction should not be used as a method of preparation of chloromethane because chloromethane also reacts with chlorine. The hydrogen atoms are substituted by chlorine atoms, one at a time, until all the hydrogen atoms have been replaced by chlorine atoms. Hydrogen chloride gas is produced at each stage.

$$CH_3Cl(g) + Cl_2(g) \rightarrow \underset{\text{dichloromethane}}{CH_2Cl_2(l)} + HCl(g)$$

$$CH_2Cl_2(l) + Cl_2(g) \rightarrow \underset{\text{trichloromethane}}{CHCl_3(l)} + HCl(g)$$

$$CHCl_3(l) + Cl_2(g) \rightarrow \underset{\text{tetrachloromethane}}{CHCl_4(l)} + HCl(g)$$

Similar reactions occur with other alkanes and chlorine.

> The reaction with chlorine is also referred to as a **photochemical reaction** – a chemical reaction initiated by light energy.
>
> In this reaction, ultraviolet light provides the activation energy, E_a, without which the molecules cannot collide successfully to form the products.

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12.3 Alkenes

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Alkenes are members of a homologous series.

- The name of each alkene ends in *-ene*.
- They have the general formula C_nH_{2n}.
- The functional group is C=C, i.e. two carbon atoms are held together with a double covalent bond. As a C=C group must be present in all alkenes, there is no alkene with one carbon atom only.
- They are **unsaturated hydrocarbons**, which means they contain at least one carbon–carbon double bond or carbon–carbon triple bond.

▼ **Table 12.5 Unbranched alkenes**

Number of carbon atoms	Molecular formula	Name	Structural formula
2	C_2H_4	Ethene	$CH_2{=}CH_2$
3	C_3H_6	Propene	$CH_3CH{=}CH_2$
4	C_4H_8	But-1-ene	$CH_3CH_2CH{=}CH_2$
4	C_4H_8	But-2-ene	$CH_3CH{=}CHCH_3$

Manufacture of alkenes

Alkenes are manufactured by **cracking** long-chain alkanes obtained from petroleum. This is a type of decomposition reaction in which carbon–carbon bonds break to form smaller molecules. Cracking requires either heat (thermal cracking) or a catalyst (catalytic cracking).

When a long-chain alkane is cracked, different molecules of the alkane may break in different places to give a mixture of products which can be separated by fractional distillation.

The mixture is likely to include:

- short-chain alkenes, used for the production of polymers and organic chemicals
- alkanes containing between 5 and 10 carbon atoms, used as fuels for petrol engines
- hydrogen, used to manufacture ammonia

For example, $C_{14}H_{30}$ molecules could crack into octane and propene:

$$C_{14}H_{30} \rightarrow C_8H_{18} + 2C_3H_6$$

or into ethene, propene and hydrogen:

$$C_{14}H_{30} \rightarrow 4C_2H_4 + 2C_3H_6 + H_2$$

(You will not be asked to predict the products of cracking without being given further information.)

Structural isomerism in alkenes

There is only one possible structure for the alkenes containing two and three carbon atoms: ethene, $CH_2{=}CH_2$, and propene, $CH_3CH{=}CH_2$.

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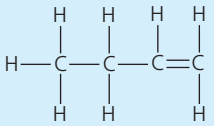
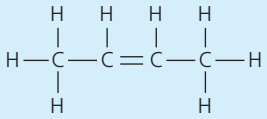
With four carbon atoms (C_4H_8), there are two unbranched alkenes because the double bond can be in two different positions in the carbon chain (Table 12.6).

▼ **Table 12.6 Structural isomerism in butene**

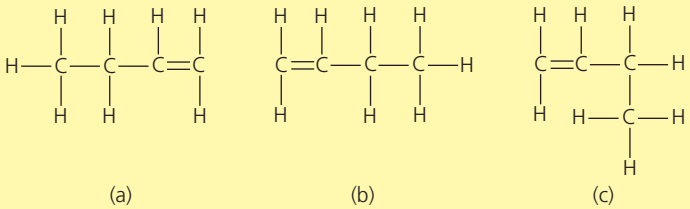
Molecular formula	C_4H_8	C_4H_8
Displayed formula		
Structural formula	$CH_3CH_2CH{=}CH_2$	$CH_3CH{=}CHCH_3$
Name	But-1-ene	But-2-ene

The number 1 in but-1-ene means that the double bond is between carbon atoms 1 and 2.

All the molecules in Figure 12.2 are but-1-ene, just drawn in different ways.

- (b) is the same as (a), only it is drawn back to front.
- (c) is also but-1-ene but the chain is bent.
- The double bond is between the first two carbon atoms in all three cases.

The number 2 in but-2-ene means that the double bond is between carbon atoms 2 and 3. Isomers must be different molecules, not the same molecule drawn differently.



▲ **Figure 12.2 Three ways of drawing the structural formula of but-1-ene**

12.4 Reactions of alkenes

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Alkenes are more reactive than alkanes because it takes less energy to convert double bonds into single bonds than to break single bonds.

Alkenes do not usually undergo substitution reactions because it is possible to add atoms to the molecules without first removing atoms. Instead, they undergo **addition reactions**, in which two molecules join together to make only one molecule.

In the addition reactions of alkenes, the double bond becomes a single bond and an atom or group of atoms joins on to each of the carbon atoms that formed the double bond.

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▲ **Figure 12.3 An addition reaction**

With hydrogen

If ethene and hydrogen are passed over a nickel catalyst at 200°C, the product is ethane:

$$CH_2{=}CH_2(g) + H_2(g) \rightarrow CH_3CH_3(g)$$

With steam

Ethene can be made to react with steam to produce ethanol using a:

- catalyst of phosphoric acid (H_3PO_4)
- temperature of 300°C
- pressure of 60 atmospheres

The equation for the reaction is:

$$CH_2{=}CH_2(g) + H_2O(g) \rightarrow CH_3CH_2OH(g)$$

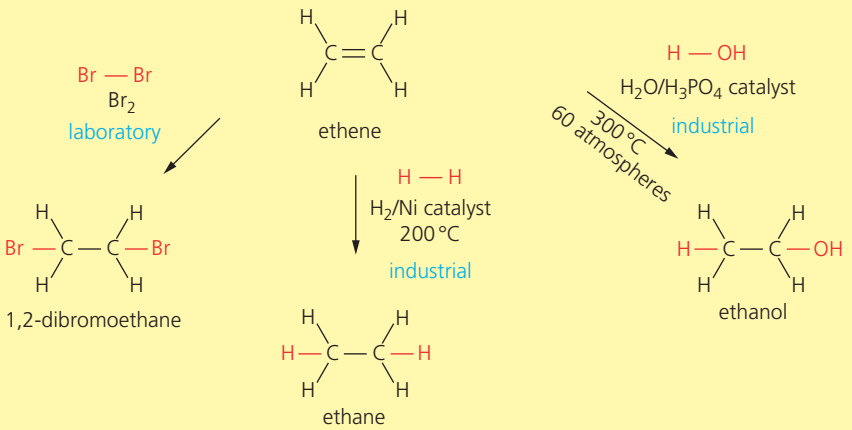
This reaction, known as the catalytic addition of steam to ethene (or hydration of ethene), is used to manufacture ethanol industrially (see Chapter 13).

With bromine

If the element bromine ($Br_2(l)$) or aqueous bromine ($Br_2(aq)$) is added to any alkene, an addition reaction occurs. If the alkene is ethene, the product is 1,2-dibromoethane:

$$CH_2{=}CH_2(g) + Br_2(l) \rightarrow CH_2BrCH_2Br(l)$$

The reactions of ethene are summarised in Figure 12.4.



▲ **Figure 12.4 The reactions of ethene**

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Skills

Testing for unsaturation

Aqueous bromine (bromine water, $Br_2(aq)$) can be used to distinguish between saturated and unsaturated substances (Table 12.7).

▼ **Table 12.7 Using bromine water to test for saturated and unsaturated substances**

	Saturated substance	Unsaturated substance
Effect of adding aqueous bromine	No change (aqueous bromine remains pale brown)	Aqueous bromine changes from pale brown to colourless

12.5 Polymers

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Polymers are large molecules made when small molecules called **monomers** take part in a reaction known as **polymerisation**.

Proteins are natural polymers that have a fixed size. Synthetic polymer molecules have no definite size.

There are two types of polymerisation reactions: addition polymerisation, discussed below, and condensation polymerisation, covered in Chapter 13.

Addition polymerisation

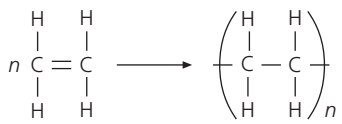
If alkenes, such as ethene, are treated to conditions of high temperature and high pressure in the presence of a suitable catalyst, the double bonds become single bonds, making more electrons available for the carbon atoms to join together.

This happens to thousands of ethene molecules, which join together to form one long-chain molecule.

- The ethene molecules are the monomers.
- The polymer is poly(ethene). Its commercial name is polythene.

This type of reaction is known as **addition polymerisation** because the monomers join together without the removal of any atoms. As in other addition reactions of alkenes, there is only one product.

The equation for the polymerisation of ethene is shown in Figure 12.5, where *n* represents a number larger than 10 000.



ethene (monomer) poly(ethene) (polymer)

▲ **Figure 12.5 Formation of poly(ethene)**

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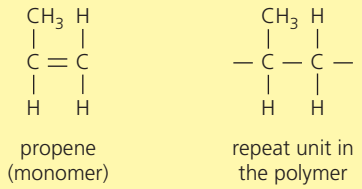
Other examples of addition polymerisation

Theoretically, any molecule with a carbon–carbon double bond can form an addition polymer.

The chemical name of a polymer is always the same as the name of the monomer with the prefix *poly-* added.

Propene, $CH_3CH{=}CH_2$, undergoes addition polymerisation to form poly(propene).



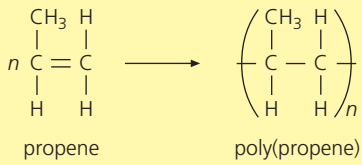
▲ **Figure 12.6 Propene monomer and polymer unit**

If you are asked to write an equation for this reaction, it is better to use the displayed formula than the structural formula.

The following formula is *incorrect*:

$$nCH_3CH{=}CH_2 \rightarrow {-}(CH_3CH{-}CH_2){-}_n$$

The carbon atom in the CH_3 group has five bonds and the carbon atom in the CH group has three bonds. A mistake like this is easier to spot if you first draw the monomer, as in Figure 12.7. Then change the double bond to a single bond and draw extension bonds on each side to show that the polymer extends in both directions. The correct equation is shown in Figure 12.7.



▲ **Figure 12.7. Formation of poly(propene)**

Plastics

Plastics are made from polymers.

- A polymer is an individual molecule, such as a poly(ethene) molecule.
- A plastic is a commercially useful material that can be made from one polymer molecule or several different polymer molecules to create objects such as buckets.

Plastics can be partially organic or fully synthetic.

Environmental challenges

Disposal

In many countries, household waste contains large quantities of plastic objects. These objects are often disposed of by:

- burying them in landfill sites
- incineration (burning)

Both of these methods contribute significantly to environmental pollution.

- Plastics buried in landfill sites remain in the environment and take up large amounts of space.

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- Incineration can lead to the production of toxic gases, such as carbon monoxide, and acidic gases, such as hydrogen chloride, which contribute to acid rain.

Attempts to overcome these problems include:

- development of biodegradable plastics (those that break down in the environment as a result of bacterial activity)
- development of photodegradable plastics (which break down in sunlight)
- sorting and recycling schemes

Accumulation in oceans

The accumulation of plastics in water sources, including oceans, was referred to in Chapter 11 (Table 11.2).

Attempts to overcome this problem include:

- decrease in manufacture of single-use plastic goods
- using materials other than plastics

Revision activity

Make a flash card for each section in this (or/and any other) chapter. Draw a picture on one side and write the key points on the other side. Shuffle the cards and sort them into groups of linked ideas. Then try doing this by looking at the pictures only, or pick a couple of cards at random and challenge yourself or a friend to use a sequence of correct statements to link the two concepts.

Sample questions

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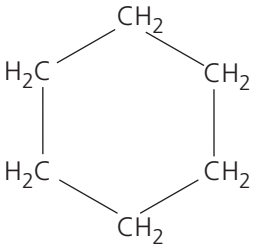
1 The structural formula of cyclohexane is shown below.



a The name gives information about the molecule.
Cyclo means that the atoms are joined in a ring.
State the meaning of:
i *hex* [1]
ii *-ane* [1]

b Give the:
i molecular formula [1]
ii displayed formula [1]
iii empirical formula [1]
of cyclohexane.

c **i** Deduce the general formula of the homologous series of compounds of which cyclohexane is a member. [1]
ii Name a homologous series that has the same general formula as that given in your answer to **(c)(i)**. [1]

d State the observations, if any, that you would expect if cyclohexane was added to aqueous bromine. Explain how you made your decision. [2]

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Student's answers

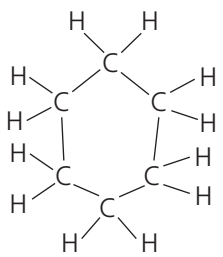
a i six
ii alkane
b i CH_2
ii $(CH_2)_6$
iii C_6H_{12}
c i CNH2N
ii alkenes
d Bromine changes from orange to colourless.

Teacher's comments

a i Six is correct but the student should have explained that six is the number of carbon atoms in one molecule.
ii A better answer would be that all the bonds are single bonds.
b The student mixed up the empirical and molecular formulae in **(i)** and **(iii)**. The displayed formula should show all the atoms and all the bonds in one molecule of cyclohexane.
c i The letter *n* should be lower case and subscript and the 2 should be subscript.
ii Alkenes is the correct answer.
d The bromine does not change colour because cyclohexane does not contain a C=C bond.

Correct answers

a i contains six carbon atoms in one molecule
ii contains single bonds only
b i C_6H_{12}
ii



iii CH_2
c i C_nH_{2n}
ii alkenes
d No colour change because cyclohexane does not contain any –C=C– bonds OR cyclohexane is saturated.

2 a Give the:
i general formula
ii molecular formula
iii structural formula
iv empirical formula
of butane.
b Name the structural isomer of butane. Give its displayed formula.

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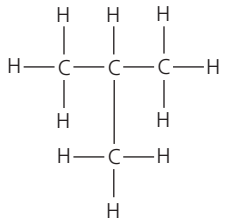
c Write the equation for the reaction between 1 molecule of the alkane with 5 carbon atoms and 1 molecule of chlorine in the presence of ultraviolet light.

Student's answers

a **i** C_nH_{2n+2}
ii C_2H_5
iii $CH_3CH_2CH_2CH_3$
iv C_4H_{10}
b 2-methylpropane



c $C_5H_{12} + 2Cl_2 \rightarrow C_5H_{10}Cl_2 + 2HCl$

Teacher's comments

a The student has the molecular formula and empirical formula the wrong way round:
- **ii** The molecular formula should show all the atoms in one molecule and give no information about how the atoms are bonded together.
- **iii** The structural formula correctly shows how atoms are arranged into groups within the molecule.
- **iv** The empirical formula should show the smallest whole number ratio of atoms of each element in the molecule.

b The student's answer is completely correct – the displayed formula shows all the atoms and all the bonds.
c The student used 2 molecules of chlorine instead of 1 molecule. If the question had referred to 2 molecules of chlorine, this would be the correct answer.

Correct answers

a **i** The general formula of all alkanes is C_nH_{2n+2}.
ii The molecular formula of butane is C_4H_{10}.
iii The structural formula of butane is $CH_3CH_2CH_2CH_3$.
iv The empirical formula of butane is C_2H_5.
b See student's answer. Note that methylpropane is an alternative name.
c Using C_nH_{2n+2} (where $n = 5$), the formula of the alkane is C_5H_{12}.
1 chlorine atom replaces 1 hydrogen atom in the substitution reaction.
The inorganic product is hydrogen chloride:

$$C_5H_{12} + Cl_2 \rightarrow C_5H_{11}Cl + HCl$$

3 Write an equation for the cracking of hexane into an alkane and an alkene, both having the same number of carbon atoms.

Student's answer

$C_6H_{16} \rightarrow C_3H_8 + C_3H_6 + H_2$

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Teacher's comment

The ending *–ane* indicates that hexane is an alkane and the general formula C_nH_{2n+2} should be used to deduce its formula. *Hex-* indicates that $n = 6$. The student begins with the incorrect formula for hexane, which makes it impossible to achieve the correct answer. H_2 is added as an attempt to 'balance' the equation.

Correct answer

Both of the following are acceptable answers because both produce an alkane and an alkene with the same number of carbon atoms:

$$C_6H_{14} \rightarrow C_3H_8 + C_3H_6$$

$$C_6H_{14} \rightarrow C_2H_6 + 2C_2H_4$$

Neither answer is more correct than the other.

If the question had specified a 1:1 mole ratio of the products, only the following would be correct:

$$C_6H_{14} \rightarrow C_3H_8 + C_3H_6$$

Exam-style questions

1 Use the following list of organic compounds to answer the questions that follow.

ethane **ethene** **methane** **nylon** **poly(ethene)**

Each substance can be used once, more than once or not at all.
For some question, you need to name only one substance. For others, there is more than one answer required.
Give the name of the substance or substances that:

- **a** are unsaturated [1]
- **b** are alkanes [2]
- **c** are formed by addition polymerisation [1]
- **d** contain a carbon–carbon double bond [1]
- **e** can act as a monomer [1]
- **f** are members of the same homologous series [2]
- **g** can be formed by hydrogenation of an alkene [1]

[Total: 9]

2 Coal gas is made by heating coal in the absence of air.
The gases listed below are the main constituents of coal gas.

carbon dioxide **carbon monoxide** **ethene** **hydrogen** **methane** **nitrogen**

- **a** Name the two gases that are hydrocarbons. [2]
- **b** Name the gas that is an alkane. [1]
- **c** Draw the displayed formula of a molecule of ethene. [1]
- **d** Describe how aqueous bromine can be used to distinguish between ethene and ethane. [2]
- **e** Name the two gases that are greenhouse gases. [2]
- **f** Ethene molecules join together to form poly(ethene).
  - **i** State the name given to this type of reaction. [1]
  - **ii** Which one of the following words describes the ethene molecules in this reaction? [1]

    elements **mixtures** **monomers** **polymers** [Total: 10]

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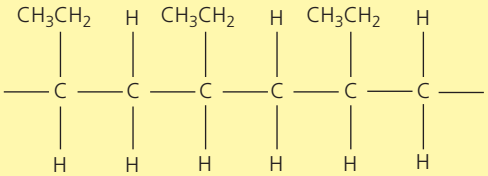
3 Tetrachloromethane, CCl_4, is a compound that is inert to most chemical reagents.
It can be produced by reacting carbon disulfide, CS_2, with chlorine in the presence of a catalyst. CCl_4 and S_2Cl_2 are the only products.

a State what is meant by the following terms:
- **i** compound [1]
- **ii** inert [1]
- **iii** catalyst. [2]

b Write a chemical equation for the reaction between carbon disulfide and chlorine. [2]

[Total: 6]

4 The diagram below shows part of a polymer.



a State the type of polymer that is shown. [1]

b Draw a circle around one repeat unit of the polymer. [1]

c **i** Draw the displayed formula of the monomer. Show all the atoms and all the bonds. [1]

ii Name the monomer. [1]

[Total: 4]

5 Draw the structures, showing all the atoms and all the bonds, of two different unbranched alkenes with the molecular formula C_5H_{10}. You are not expected to name the alkenes. [Total: 2]

6 Ethane reacts with chlorine in a substitution reaction.

a Under what condition does the reaction take place? [1]

b Name the organic product formed when ethane and chlorine react in a 1:1 mole ratio. [1]

c If an excess of chlorine is used, give the molecular formula of one other organic product that could form. [1]

[Total: 3]

7 Propene reacts with:
- **i** bromine
- **ii** hydrogen
- **iii** steam

a Name the type of reaction that occurs in all three cases. [1]

b State the observation you would expect to see in reaction **(i)** if excess propene is used. [2]

c Write down the molecular formulae of the products that form in reactions **(i)**, **(ii)** and **(iii)**. [3]

d Name the catalyst used in **(ii)**. [1]

[Total: 7]

8 Alkanes are converted into alkenes by cracking.

a Give the molecular formula of the alkane that contains nine carbon atoms. [1]

b Draw the structure and give the name of an alkene with four carbon atoms. [2]

c Write an equation for the cracking of octane, C_8H_{18}, into:
- **i** an alkane and an alkene formed in the mole ratio 1:2 [2]
- **ii** hydrogen and two other products [2]

[Total: 7]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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13 Organic chemistry 2

Key objectives

By the end of this section, you should be able to:

- write and interpret the general formulae of alcohols and carboxylic acids
- state the type of compound present, given a chemical name ending in -ol or -oic acid, or from a molecular formula or displayed formula
- name and draw the displayed formulae of ethanol, ethanoic acid and the products of the reactions referred to in this chapter

- name and draw the structural and displayed formulae of unbranched alcohols and carboxylic acids containing up to four carbon atoms per molecule

- describe the manufacture of ethanol by fermentation and by catalytic addition of steam to ethene

- describe the advantages and disadvantages of the manufacture of ethanol by fermentation and by catalytic addition of steam to ethene

- describe the combustion of ethanol
- state the uses of ethanol
- describe the reactions of carboxylic acids with metals, bases and carbonates

- describe the formation of ethanoic acid by oxidation of ethanol
- describe the reaction of a carboxylic acid with an alcohol using an acid catalyst to form an ester
- name and draw the displayed formulae of the unbranched esters which can be made from unbranched alcohols and carboxylic acids each containing up to four carbon atoms
- describe the difference between addition polymerisation and condensation polymerisation
- deduce the structure or repeat unit of a polyamide or polyester from given monomers and vice versa
- describe and draw the structure of nylon and PET
- state that PET can be converted back into monomers and repolymerised
- describe proteins as natural polyamides formed from amino acids
- describe and draw the structure of proteins

Key terms

REVISED

Term	Definition
Combustion	A chemical reaction in which a substance reacts rapidly with oxygen, producing heat and light.
Condensation polymer	A polymer formed by a condensation reaction (a reaction in which a simple molecule, such as water, is produced during polymerisation).
Fermentation	A series of biological reactions, catalysed by the enzymes in yeast.
Functional group	An atom or group of atoms responsible for the characteristic chemical reactions of an organic compound.

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13.1 Functional groups

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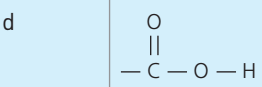
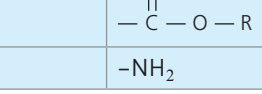
Examples of **functional groups** are shown in Table 13.1.

▼ **Table 13.1 Functional groups**

Homologous series	Functional group
Alkene	–C=C–
Alcohol	–O–H
Carboxylic acid	
Ester	
Amine	$-NH_2$

13.2 Alcohols

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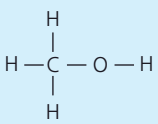
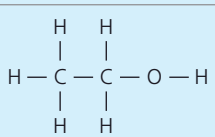
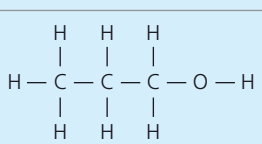
Formulae and names of alcohols

Alcohols are members of a homologous series.

- The names of alcohols all end in *-ol*.
- The general formula of an alcohol is $C_nH_{2n+1}OH$.
- Alcohols contain the –O–H functional group.
- The structural formula of ethanol can be written as either CH_3CH_2OH or C_2H_5OH.

- Alcohols with more than two carbon atoms have unbranched structural isomers because the O–H group can be in different positions on the carbon chain. A number is used to indicate the position of the O–H group (see Table 13.1).

▼ **Table 13.2 Alcohols**

Number of carbon atoms	Displayed formula	Structural formula	Name
1		CH_3OH	Methanol
2		CH_3CH_2OH or C_2H_5OH	Ethanol
3		$CH_3CH_2CH_2OH$	Propan-1-ol

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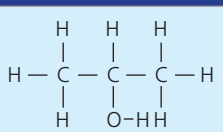
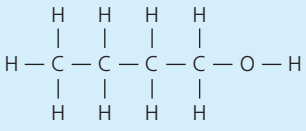
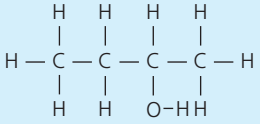
Number of carbon atoms	Displayed formula	Structural formula	Name
4		$CH_3CH(OH)CH_3$ or $CH_3CH(OH)CH_3$	Propan-2-ol
5		$CH_3CH_2CH_2CH_2OH$	Butan-1-ol
6		$CH_3CH_2CH(OH)CH_3$ or $CH_3CH_2CH(OH)CH_3$	Butan-2-ol

Manufacture of ethanol

Ethanol is manufactured on a large scale by **fermentation** of carbohydrates and **catalytic addition** of steam to ethene.

Fermentation of carbohydrates

Carbohydrates, such as sugar, are broken down by enzymes in yeast to produce glucose, $C_6H_{12}O_6$. The enzymes also catalyse the breakdown of glucose to form ethanol and carbon dioxide. The reaction occurs at a temperature of 37°C and is carried out in the absence of oxygen:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

When the concentration of ethanol reaches 14%, it kills the yeast. The yeast cells are removed by filtration and the ethanol is purified by fractional distillation (see Chapter 14).

Catalytic addition of steam to ethene

Ethene reacts with steam to produce ethanol, as described in Section 12.4:

$$C_2H_4 + H_2O \rightarrow C_2H_5OH$$

The advantages and disadvantages of the two processes are shown in Table 13.3.

▼ **Table 13.3 Advantages and disadvantages of methods of producing ethanol**

	Fermentation	Catalytic addition
Advantages	Uses carbohydrates from plants, which are a renewable resource	There is only one product in the reaction, which means there is no waste
	Requires a temperature of 37°C, which means energy costs are low	A continuous flow process is used, which is efficient
Disadvantages	A batch process is used, which is inefficient	Uses ethene from petroleum, which is a non-renewable resource
	Land which could be used to grow plants for food is used for ethanol production	Requires a temperature of 300°C, which means energy costs are high

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Uses of ethanol

Ethanol is used as:

- a solvent
- a fuel in spirit camping stoves, where it undergoes complete **combustion** to produce carbon dioxide and water:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

13.3 Carboxylic acids

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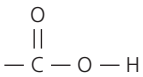
Carboxylic acids are members of a homologous series.

- The names of carboxylic acids all end in *-oic acid*.
- The general formula of an carboxylic acid is $C_nH_{2n+1}COOH$.
- The functional group in carboxylic acids is –COOH, which can also be written as $–CO_2H$. This is displayed as shown in Figure 13.1.

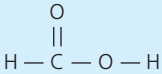
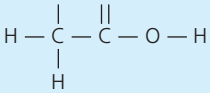
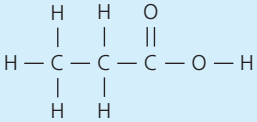
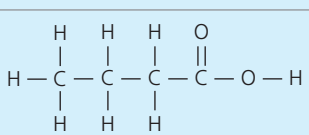


▲ **Figure 13.1 Carboxylic acid functional group**

- The formulae of most organic compounds begin with a C atom. However, make sure you remember that the structural formula for methanoic acid is written as HCOOH.

Unbranched carboxylic acids with up to four carbon atoms are shown in Table 13.4.

▼ **Table 13.4 Carboxylic acids**

Number of carbon atoms	Displayed formula	Structural formula	Name
1		HCOOH	Methanoic acid
2		CH_3COOH	Ethanoic acid
3		CH_3CH_2COOH	Propanoic acid
4		$CH_3CH_2CH_2COOH$	Butanoic acid

When drawing the displayed formulae for carboxylic acids and alcohols, remember to show the bond between O and H atoms, i.e. draw –O–H rather than –OH.

Making ethanoic acid

Ethanoic acid is formed in the laboratory by oxidation of ethanol using acidified aqueous potassium manganate(VII), which acts as an oxidising agent.

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Skills

Laboratory conversion of ethanol to ethanoic acid

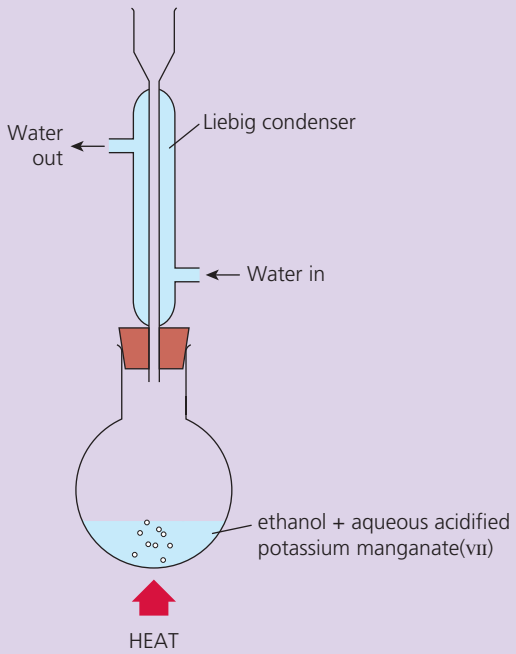
Ethanol is converted into ethanoic acid using the apparatus shown in Figure 13.2. The technique used is known as **heating under reflux**.

The reaction mixture is heated for at least 30 minutes. The organic vapours pass into the Liebig condenser, where they condense and fall back into the reaction vessel. This prevents any loss of ethanoic acid produced and allows heating for as long as is necessary to achieve a good yield of ethanoic acid.

After heating under reflux, the ethanoic acid is separated from the rest of the reaction mixture by distillation.

A simplified version of the equation, which represents oxygen from the oxidising agent as [O], is:

$$CH_3CH_2OH + 2[O] \rightarrow CH_3COOH + H_2O$$



▲ **Figure 13.2 Making ethanoic acid**

Bacterial oxidation of ethanol

Ethanol can also be oxidised to ethanoic acid using oxygen in the air as the oxidising agent:

$$CH_3CH_2OH + O_2 \rightarrow CH_3COOH + H_2O$$

The reaction is catalysed by enzymes in bacteria and is used in the production of vinegar.

Reactions of aqueous ethanoic acid

Ethanoic acid, CH_3COOH, is a typical **weak acid**. It reacts with metals, bases and carbonates to produce salts (see Chapter 8). The salts are called **ethanoates** and contain the ethanoate ion, CH_3COO^-.

With metals

Metals above hydrogen in the reactivity series, e.g. magnesium, react with dilute ethanoic acid. The solid disappears and an aqueous solution forms. The colour of the aqueous solution that forms depends on the metal present. Bubbles are seen because hydrogen gas is produced.

$$2CH_3COOH(aq) + Mg(s) \rightarrow (CH_3COO)_2Mg(aq) + H_2(g)$$

magnesium ethanoate

With bases

Aqueous alkalis (soluble bases), e.g. aqueous sodium hydroxide, neutralise dilute ethanoic acid. There are *no* observations unless an indicator is present.

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

sodium ethanoate

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Solid insoluble bases, e.g. copper(II) oxide, disappear when added to dilute ethanoic acid and form an aqueous solution. The colour of the aqueous solution that forms depends on the metal present in the base.

$$2CH_3COOH(aq) + CuO(s) \rightarrow (CH_3COO)_2Cu(aq) + H_2O(l)$$

copper(II) ethanoate

With carbonates

Carbonates react with dilute ethanoic acid. Solid carbonates, e.g. calcium carbonate, disappear and an aqueous solution is formed. The colour of the aqueous solution that forms depends on the metal present in the carbonate. Bubbles are seen because carbon dioxide gas is formed.

$$2CH_3COOH(aq) + CaCO_3(s) \rightarrow (CH_3COO)_2Ca(aq) + CO_2(g) + H_2O(l)$$

calcium ethanoate

Notice that the symbol for the metal appears at the end of the formula for each of these salts rather than at the beginning, as it does for inorganic salts.

Revision activity

Figure 12.4 (page 127) shows the reactions of ethene in a diagram. Create similar diagrams for ethane, ethanol and ethanoic acid.

13.4 Esters

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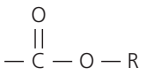
Esters are sweet-smelling liquids.

- The names of esters all end in *-oate*, as with salts of carboxylic acids.
- Esters have a general formula of $C_nH_{2n}O_2$.
- The functional group in esters is –COOR, which can also be written as $–CO_2R$ and is displayed as shown in Figure 13.3. R represents a group containing carbon and hydrogen atoms.



▲ **Figure 13.3 Ester functional group**

Esterification

Esterification is the name given to the reaction between a carboxylic acid and an alcohol.

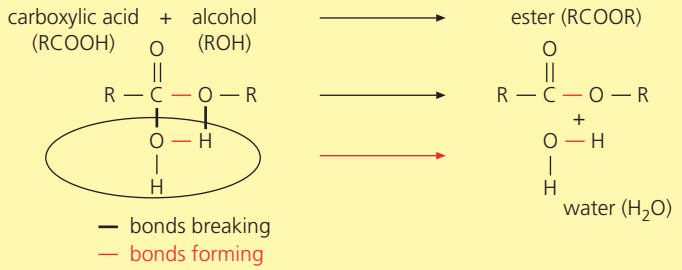
The alcohol and carboxylic acid are heated with a catalyst of concentrated sulfuric acid. The general equation for this in words is:

carboxylic acid + alcohol → ester + water

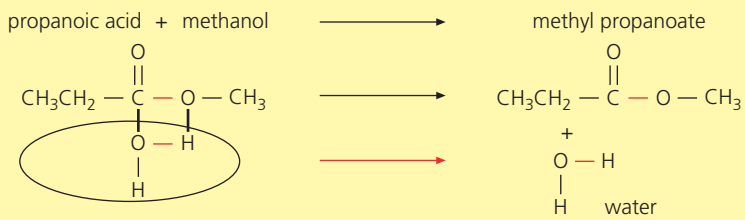
The molecules can be represented as:

$$RCOOH + ROH \rightarrow RCOOR + H_2O$$

The reaction occurs as shown in Figure 13.4.



Example



▲ **Figure 13.4 The formation of an ester**

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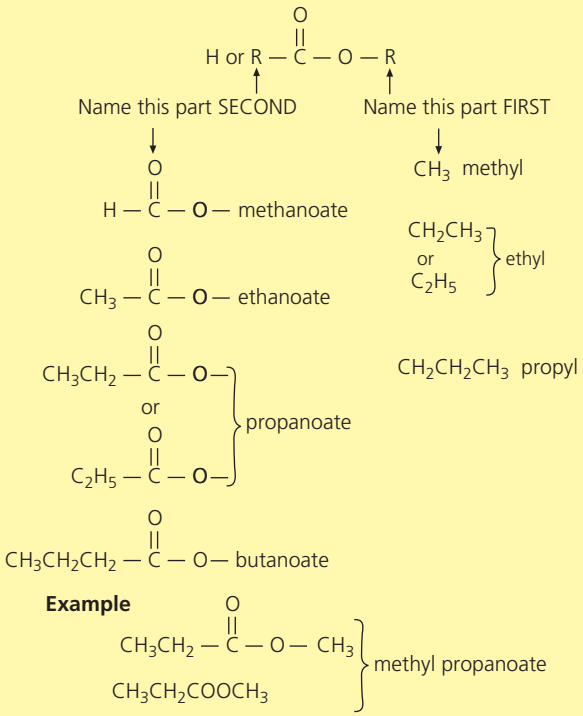
Another example is:

$$CH_3COOH(l) + C_2H_5OH(l) \rightarrow CH_3COOC_2H_5(l) + H_2O(l)$$

ethanoic acid ethanol ethyl ethanoate water

Naming esters

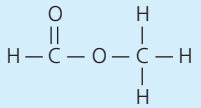
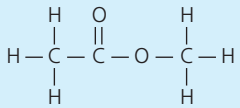
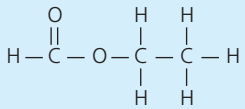
Naming esters is unlike naming any other organic molecules we have met so far. The formula is divided into two and each part is named according to the number of carbon atoms it contains, as shown in Figure 13.5. There are no esters with one carbon atom.



▲ **Figure 13.5 Naming esters**

The names and formulae of unbranched esters are shown Table 13.5. Notice that the formula of methyl methanoate is written $HCOOCH_3$. As with methanoic acid, from which it is made, the structural formula begins with an H. The same thing applies to salts made from methanoic acid.

▼ **Table 13.5 Names and formulae of unbranched esters**

Number of carbon atoms	Molecular formula of ester	Displayed formula	Structural formula	Name of ester	Made from	
					Carboxylic acid	Alcohol
2	$C_2H_4O_2$		$HCOOCH_3$	Methyl methanoate	Methanoic acid	Methanol
3	$C_3H_6O_2$		CH_3COOCH_3	Methyl ethanoate	Ethanoic acid	Methanol
3	$C_3H_6O_2$		$HCOOCH_2CH_3$	Ethyl methanoate	Methanoic acid	Ethanol

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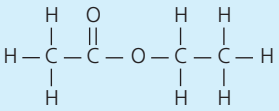
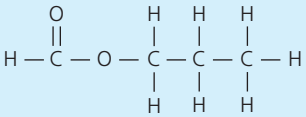
Number of carbon atoms	Molecular formula of ester	Displayed formula	Structural formula	Name of ester	Made from	
					Carboxylic acid	Alcohol
4	$C_4H_8O_2$		$CH_3COOCH_2CH_3$	Ethyl ethanoate	Ethanoic acid	Ethanol
4	$C_4H_8O_2$		$HCOOCH_2CH_2CH_3$	Propyl methanoate	Methanoic acid	Propan-1-ol

13.5 Condensation polymerisation

REVISED

A condensation reaction is a reaction in which two molecules join together and a simple molecule, such as water, is removed at the same time. Esterification is an example of a condensation reaction.

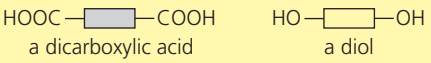
Condensation polymers are formed from monomers with two functional groups each. Examples of such functional groups are –OH, –COOH and $–NH_2$.

The repeat unit of a condensation polymer contains what is left of the two monomers in the polymer after polymerisation has taken place.

Polyesters and polyamides are examples of condensation polymers.

Polyesters

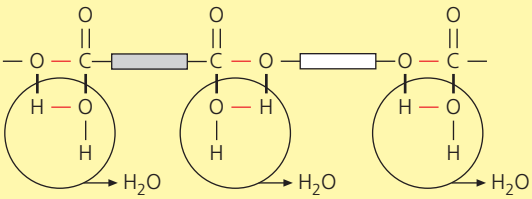
Polyesters can be made from dicarboxylic acids (molecules with two –COOH groups) and diols (molecules with two –OH groups). These monomers can be represented as shown in Figure 13.6.



▲ **Figure 13.6 Block representation of complex molecules**

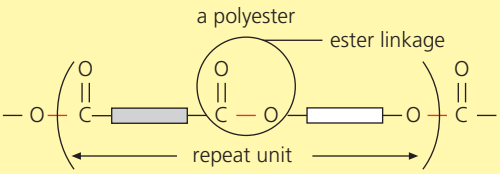
The polymerisation occurs by the removal of a molecule of water when a –COOH group and an –OH group react, as shown in Figure 13.7.

The diagram shows the formation of one repeat unit of the polyester. However, because there are –COOH groups and –OH groups at both ends of the monomers, more linkages can form and the polymer chain can grow in both directions.



— bonds breaking

— bonds forming



▲ **Figure 13.7 Condensation polymerisation**

Revision activity

Make a set of 10 cards with names of organic compounds from the syllabus, e.g. but-2-ene, and another set of 10 cards with their structural formulae, e.g. $CH_3CH{=}CHCH_3$. One player puts down a card from the set with names and the second player matches the name with the card showing the correct formula. You can decide on a scoring system and swap cards after one round.

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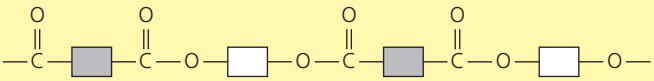
PET is a polyester used in clothing manufacture. It is made from the dicarboxylic acid HOOC–C_6H_4–COOH and the diol HO–C_2H_4–OH.

The structure of part of the PET molecule is shown in Figure 13.8.

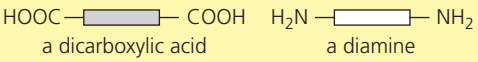


▲ **Figure 13.8 PET**

PET can be converted back into its monomers and repolymerised. This means that disposal is less of an environmental challenge than with some other polymers.

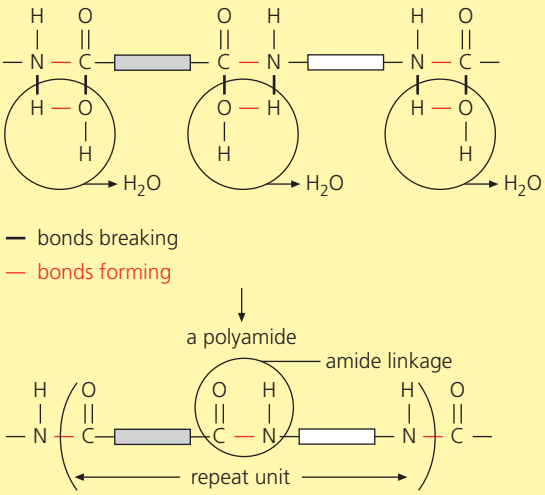
Polyamides

Polyamides can be made from dicarboxylic acids (molecules with two –COOH groups) and diamines (molecules with two –NH_2 groups). These monomers can be represented as shown in Figure 13.9.



▲ **Figure 13.9 Base units for polyamides**

The polymerisation occurs by the removal of a molecule of water when a –COOH group and a –NH_2 group react. The monomers join together as shown in Figure 13.10.

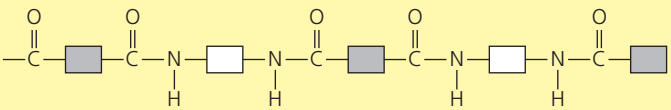


▲ **Figure 13.10 Formation of polyamides**

Once again, the –COOH groups and –NH_2 groups at each end of the monomers allow the polymer chain to grow in both directions.

Nylon is a synthetic polyamide made from the dicarboxylic acid HOOC–C_4H_8-COOH and the diamine H_2N–C_6H_{10}–NH_2.

The structure of part of a nylon molecule is shown in Figure 13.11.



▲ **Figure 13.11 Nylon**

Table 13.6 shows the differences between addition polymerisation and condensation polymerisation.

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▼ **Table 13.6 Differences between addition and condensation polymerisation**

	Addition	Condensation
Monomers	Contain a C=C double bond	Contain two reactive functional groups each, e.g. $-NH_2$, –COOH, –OH
Polymerisation	Occurs without any loss of atoms, producing only one product (the polymer)	Occurs with removal of a simple molecule, e.g. water, producing two products
Polymers	Have the same empirical formula as the monomer	Have a different empirical formula from the monomers

13.6 Natural polyamides

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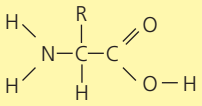
Proteins are natural polyamides which are made from amino acid monomers. There are 20 different amino acids, but all contain an $-NH_2$ (amine) and a –COOH (carboxylic acid) functional group, as shown in Figure 13.12.

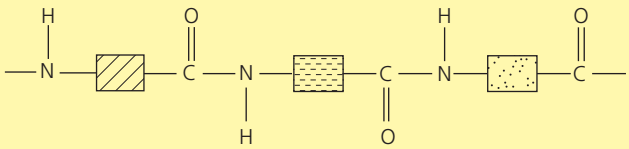


▲ **Figure 13.12 Amino acids**

The $-NH_2$ and –COOH groups react together to produce polymers which have amino acid residues in a sequence which is specific to each individual protein.

Proteins contain the same amide linkage as that present in synthetic polyamides, such as nylon, although biologists usually refer to it as a peptide linkage.

The structure of proteins can be represented as shown in Figure 13.13.



▲ **Figure 13.13 General structure of a protein**

Sample questions

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1 The following is a list of organic compounds:

pentanoic acid **hex-3-ene** **octan-2-ol** **heptane**

a Name the compound which is an:

i alkane [1]

ii alkene [1]

iii alcohol [1]

b Name the compound which contains the functional group:

i –OH [1]

ii –C=C– [1]

iii –COOH [1]

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Student's answers

a i heptane
ii hex-3-ene
iii octan-2-ol
b i octan-2-ol
ii hex-3-ene
iii pentanoic acid

Teacher's comments

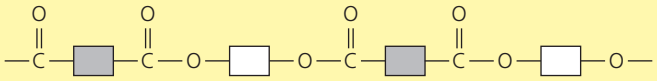
The student's answers are all correct.

Even though the compounds listed are unfamiliar, you are expected to be able to deduce which homologous series each belongs to from the ending of the name and the functional group. Remember:

a The names of:
- i alkanes end in *-ane*
- ii alkenes end in *-ene*
- iii alcohols end in *-ol*.

b
- i Alcohols contain –OH.
- ii Alkenes contain –C=C–.
- iii Carboxylic acids contain –COOH.

2 The diagram below shows part of a condensation polymer.



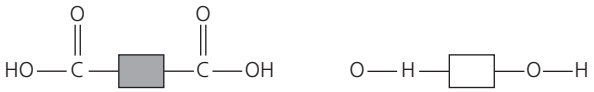
a State the name of the *type* of condensation polymer.

b Complete the diagrams below to show the functional groups in the two monomers that make the polymer. Show all the atoms and all the bonds.

Student's answers

a PET

b



Teacher's comments

a The student has ignored the word *type* and has given the name of a specific polymer.

b The student has not shown the O–H bonds in the dicarboxylic acid. In the diol, the student has drawn O–H–, giving the hydrogen two bonds when it should make only one.

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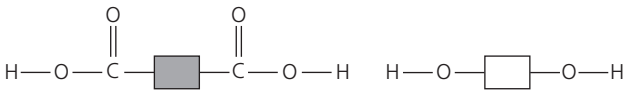
Correct answers

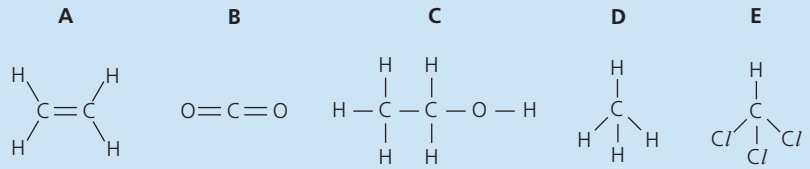
a a polyester

b



Exam-style questions

1 The diagram below shows the structures of five compounds, **A**, **B**, **C**, **D** and **E**.



Answer these questions using the letters **A**, **B**, **C**, **D** and **E**.
Each letter may be used once, more than once or not at all.
Give the letter which represents a compound that:

a is an unsaturated hydrocarbon
b is an alkane
c can be produced by catalytic addition to ethene
d is a product of complete combustion of hydrocarbons
e is produced as a waste gas from digestion in animals
f is a waste gas produced in fermentation
g decolourises aqueous bromine [Total: 7]

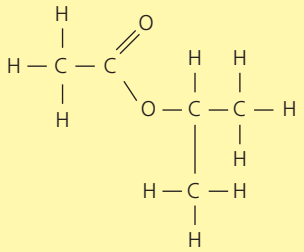
2 Dilute ethanoic acid reacts with the same substances as dilute inorganic acids to form salts as well as other products.
The solids below are added to separate samples of dilute ethanoic acid:

a zinc [4]
b magnesium carbonate [4]

For the reaction of each solid with the acid:

i name the salt produced
ii write the word equation
iii give two observations [Total: 8]

3 Three different compounds, **A**, **B** and **C**, all have the molecular formula C_3H_8O.
Compound **A** reacts with ethanoic acid to produce a compound with the structure shown below.



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Compound **B** reacts with ethanoic acid to produce a compound with the molecular formula $C_5H_{10}O_2$.
Compound **C** does not react with ethanoic acid, but it undergoes complete combustion when burned in excess oxygen.

a What is meant by the term *molecular formula*? [1]

b What is the term used to describe compounds with the same molecular formula but different structural formulae? [1]

c What is the empirical formula of the compound with the molecular formula $C_5H_{10}O_2$? [1]

d What type of reaction occurs between compound **A** and ethanoic acid? [1]

e What are the conditions that are required for compound **A** to react with ethanoic acid? [2]

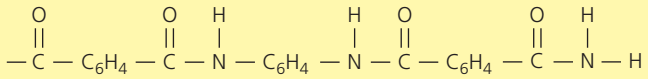
f Complete the chemical equation for the reaction occurring when compound **C** undergoes complete combustion in excess oxygen. State symbols are not required. [2]

$$C_3H_8O + O_2 \rightarrow$$

g Draw the structures of molecules **A**, **B** and **C**. Show all of the atoms and all of the bonds. [3]

[Total: 11]

4 The diagram below shows part of a polymer which is formed by condensation polymerisation.



a State the meaning of *condensation polymerisation*. [2]

b State the type of condensation polymer that is shown. [1]

c Draw a circle around one repeat unit of the polymer. [1]

d Draw a circle around the linkage in the polymer. [1]

e State the type of biological molecule containing the same linkage as the polymer. [1]

f Draw the structures of the two monomers, showing all the atoms and bonds in the functional groups. (You may leave C_6H_4 as it is written.) [2]

[Total: 8]

5 **a** Give the structural formulae and the names of the two isomeric esters with *three* carbon atoms each. [4]

b Give the structural formulae and the names of the three isomeric unbranched esters with *four* carbon atoms each. [6]

[Total: 10]

Answers available at: **www.hoddereducation.co.uk/cambridgeextras**

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14 Experimental techniques and chemical analysis

Key objectives

By the end of this section, you should be able to:

- name appropriate apparatus for the measurement of time, temperature, mass and volume
- suggest advantages and disadvantages of experimental methods and apparatus
- describe solvents, solutes, solutions, saturated solutions, residue and filtrate
- describe the use of paper chromatography to separate mixtures
- interpret simple chromatograms to identify:
  - unknown substances
  - pure and impure substances
- describe the meaning and use of R_f values and locating agents
- describe and explain methods of separation and purification, including:
  - use of a suitable solvent
  - filtration
  - crystallisation
  - simple distillation
  - fractional distillation
- suggest suitable separation and purification techniques given information about the substances involved
- identify substances and assess their purity using melting point and boiling point information
- describe tests to identify aqueous cations: aluminium, ammonium, calcium, chromium(III), copper(II), iron(II), iron(III) and zinc, including their results
- describe use of flame tests to identify cations: lithium, sodium, potassium, copper(II), barium and calcium, including their results
- describe tests to identify anions: chloride, bromide, iodide, carbonate, sulfite, sulfate and nitrate, including their results
- describe tests to identify gases: ammonia, carbon dioxide, oxygen, hydrogen, chlorine and sulfur dioxide, including their results

Key terms

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Term	Definition
Chromatography	A method of separation of mixtures of dissolved substances.
Crystallisation	The process of forming crystals from a liquid.
Filtrate	A liquid or solution that has passed through a filter paper.
Filtration	The process of separating a solid from a liquid using a filter paper which does not allow the solid to pass through.
Fractional distillation	A method of separation of a mixture of liquids with different boiling points.
Saturated solution	A solution containing the maximum concentration of a solute dissolved in the solvent at a specified temperature.
Residue	A substance that remains after evaporation, distillation or any similar process.
(Simple) distillation	The process of using evaporation and condensation to form a pure liquid from a solution.
Solute	A substance that is dissolved in a solvent.
Solution	A mixture of one or more solutes dissolved in a solvent.
Solvent	A substance that dissolves a solute to form a solution.
Substance	A general term that refers to elements, mixtures and compounds.

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14.1 Apparatus used for measurement in chemistry

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Table 14.1 shows apparatus that is appropriate for measurements of different quantities to different degrees of accuracy.

▼ **Table 14.1 Measuring apparatus**

Apparatus	Quantity measured	Further information
Stopwatch	Time	Available with different accuracies, e.g. to the nearest second or 0.1 second
Thermometer	Temperature	Available with different accuracies, e.g. to the nearest degree Celsius or 0.1 degree Celsius
Balance*	Mass	Available with different accuracies, e.g. to the nearest 0.1 gram or 0.01 gram
Burette	Volume of liquid	Usually accurate to the nearest 0.1 cm^3 Used in titrations Can be inverted and filled with water to measure gas volumes
(Volumetric) pipette	Volume of liquid	Usually only accurate to the nearest 0.1 cm^3 Used in titrations
Measuring cylinder	Volume	Usually accurate to the nearest 0.1 cm^3 Can be inverted and filled with water to measure gas volumes
Gas syringe	Volume of gas	Usually only accurate to the nearest 1.0 cm^3

* Make sure you use the correct name for each piece of apparatus, for example, *balance* rather than *weighing machine*.

14.2 Separating mixtures

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▼ **Table 14.2 Separation techniques**

Method of separation	Example of mixture that is separated with this method	Property that the method depends upon
Filtration	Muddy water	Solubility
Crystallisation	Sodium chloride solution	Solubility at different temperatures
(Simple) distillation	Sodium chloride solution	Boiling point
Fractional distillation	Ethanol and water	Boiling point
Paper chromatography	Dyes in ink	Adsorption by paper Solubility in solvent

Revision activity

Write a sentence to summarise each row of Table 14.2. Use all the key terms. When you have finished, try writing the sentences again without looking at the table.

Dissolving, filtration and crystallisation

These methods can be used in sequence to separate a mixture of two solids, one of which is soluble in a given solvent and the other of which is insoluble.

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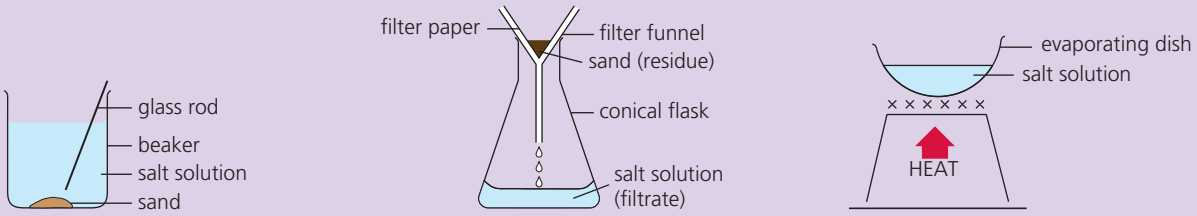
Skills

Combining separation techniques

This method can be used to separate a mixture of common salt and sand to produce pure samples of both solids.

- If the mixture is not powdered, it should be ground into a powder using a mortar and pestle. The powder is then added to water in a beaker. The common salt dissolves and the sand remains undissolved.
- The mixture is then transferred to the filtration apparatus. The sand (**residue**) remains in the filter paper and the salt solution (**filtrate**) passes through into the conical flask. This **filtration** (often spelt wrongly as *filt**e**ration*) can also be called *filtering*.
  There are other common errors often made when describing this process:
  - The words *residue* and *filtrate* are often used the wrong way round.
  - *Filtrate* is often used as an incorrect alternative to *filtered*, as in *he filtrated the solution*. The word *filtrated* does not exist.
- To obtain pure sand, distilled water should be passed through the filter paper (known as washing the residue) and then the filter paper should be removed and dried in a low oven or on a warm windowsill.
- To obtain salt crystals, the salt solution should **crystallised**:
  - First it is heated in an evaporating dish until about half the water has been removed. An alternative way of knowing when to stop heating is to dip a glass rod into the solution. If crystals form on the rod when it is removed from the solution, it is time to stop.
  - The filtrate should not be heated until *all* the water evaporates, as any water of crystallisation (see Chapter 8) would be driven off.
  - The hot **saturated salt solution** should be allowed to cool down slowly. Crystals of salt will then form.
- If there is any liquid above the crystals, it should be removed by filtration.
- The salt crystals should then be dried in a low oven or on a warm windowsill.



▲ **Figure 14.1 Dissolving, filtration and crystallisation**

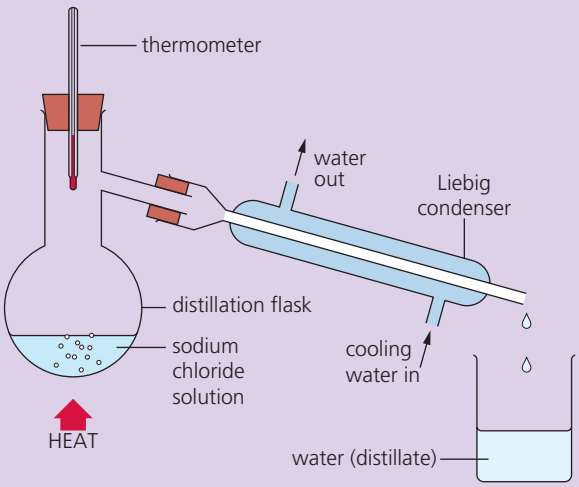
(Simple) distillation

(Simple) distillation is a method of separating a pure liquid from a solution.

Skills

Simple distillation

Pure water can be separated from a solution of sodium chloride by simple distillation using the apparatus shown in Figure 14.2.



▲ **Figure 14.2 Simple distillation**

- The flask is heated.
- The water in the sodium chloride solution evaporates and water vapour/steam enters the Liebig condenser, where it condenses as water.
- The water drips out of the end of the Liebig condenser and collects in the beaker.
- The water is pure and is referred to as **distilled water**.
- Sodium chloride does not vaporise or even melt because it has a very high melting point (see Chapter 3) and, therefore, it remains in the distillation flask.

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Fractional distillation

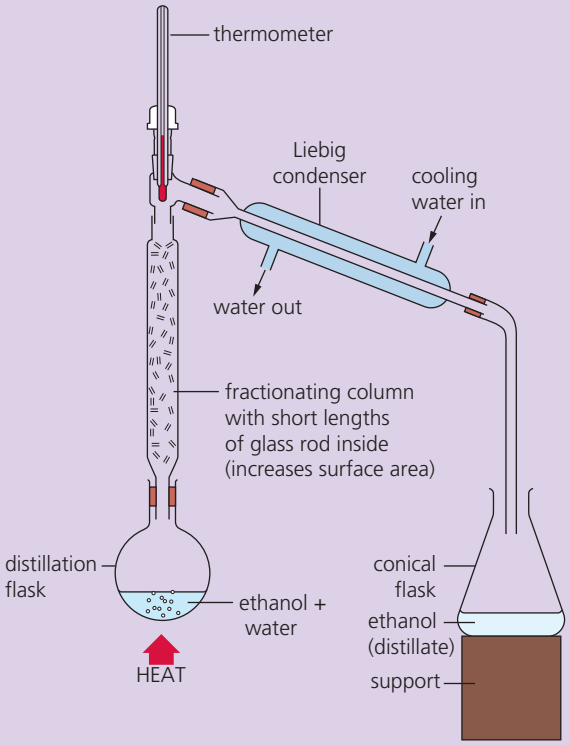
Fractional distillation is a method of separating two (or more) miscible liquids with different boiling points. It can be carried out in the laboratory or on an industrial scale, as in the fractional distillation of petroleum (see Chapter 6).

Skills

Fractional distillation

In the laboratory, ethanol and water can be separated by fractional distillation using the apparatus shown in Figure 14.3.

- Ethanol has a boiling point of 78°C and water has a boiling point of 100°C.
- The flask is heated and ethanol vapour enters the fractionating column.
- However, some water also evaporates (below its boiling point) and enters the fractionating column as water vapour/steam.
- The water vapour/steam condenses in the fractionating column and drips back down into the distillation flask.
- When the temperature reaches 78°C, the ethanol vapour reaches the top of the fractionating column and enters the Liebig condenser where it condenses.
- Finally, liquid ethanol collects as the distillate and all the water remains in the distillation flask.



▲ **Figure 14.3 Apparatus for fractional distillation**

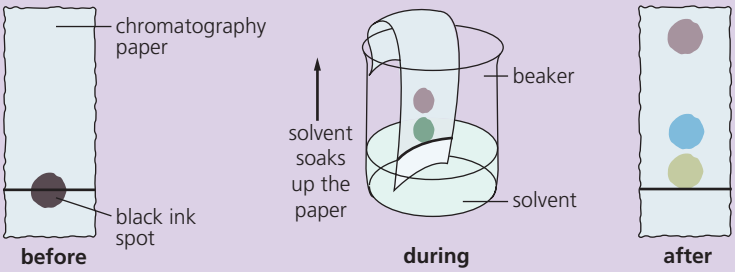
Chromatography

Chromatography can be used to separate the components of solutions which contain several dissolved substances. The substances are often coloured but may be colourless.

Skills

Chromatography as a separation technique

Paper chromatography can be used to separate the dyes in ink.



▲ **Figure 14.4 Paper chromatography**

- A spot of the ink is placed on the chromatography paper.
- The paper is placed in a suitable solvent in a beaker. If the solvent is volatile (vaporises easily), it is necessary to put a lid on the beaker to prevent the vapour from escaping.
- As the solvent rises, the dyes in the ink separate.

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Skills

Chromatography for analysis

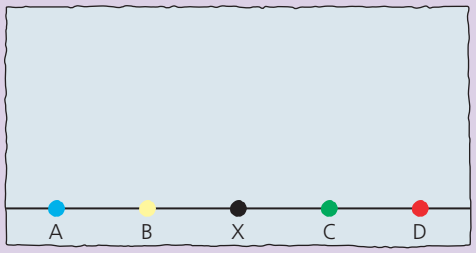
Chromatography can also be used to identify the components of a mixture as well as to separate them.

- A mixture of dyes is placed on chromatography paper in the position marked **X**, as shown in Figure 14.5.
- Four dyes whose identities are known are placed in positions marked **A**, **B**, **C** and **D**. These four dyes are referred to as **standards**.
- Chromatography is then carried out and the chromatography paper (also known as a **chromatogram**) is removed from the beaker and dried.
- The paper is then labelled to show what mixture **X** contains, as shown in Figure 14.6.

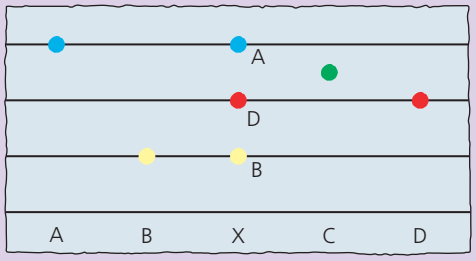
Mixture **X** is composed of three dyes because the mixture has been separated into three parts.

The three dyes are **A**, **B** and **D**. We know this because the three dyes in mixture **X** have travelled the same distances as the three standards **A**, **B** and **D** whose identities are known.

We can also conclude that mixture **X** does *not* contain dye **C** because none of the components of **X** travelled the same distance as dye **C**.



▲ **Figure 14.5 Before chromatography**



▲ **Figure 14.6 After chromatography**

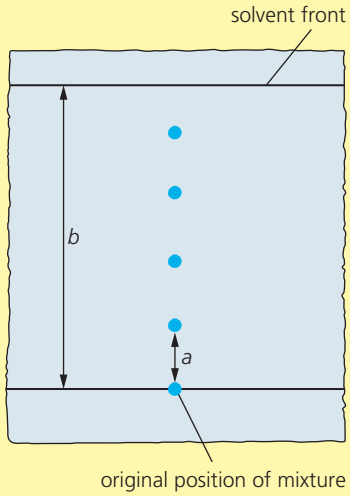
Chromatography can also be used to identify colourless substances. The experimental technique is the same, but because the components of the mixture are colourless, the spots on the chromatography paper are invisible.

After drying, the paper is sprayed with a **locating agent**, which reacts with the components of the mixture to produce coloured spots.

Instead of using standards as described above, components of a mixture can be identified by their **R_f values**.

After the chromatogram has dried, the distance that the solvent has travelled and the distance that the component of the mixture has travelled are both measured, as shown in Figure 14.7.

When the R_f value is calculated, the component of the mixture can be identified by comparison with R_f values in a data book. R_f values can be determined for all components of the mixture.



$$R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}} = \frac{a}{b}$$

▲ **Figure 14.7 Calculating R_f**

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14.3 Qualitative analysis

REVISED

Cations

Cations (positive ions) can be identified using:

- aqueous sodium hydroxide, as described in Chapter 10
- flame tests
- aqueous ammonia

Skills

Flame tests

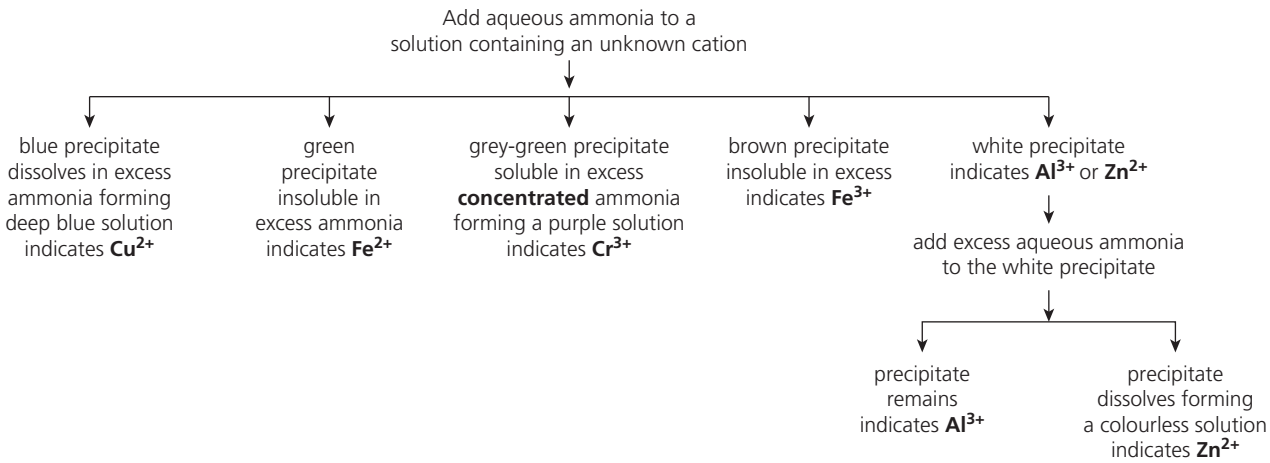
Flame tests can be carried out on solids or on aqueous solutions.

- Starting with a solid, a few drops of concentrated hydrochloric acid are added to a sample of the solid on a watch glass. Dilute hydrochloric acid can be used to avoid safety issues.
- A small amount of the mixture should then be placed on a nichrome wire.
- The nichrome wire containing some of the mixture is then placed in the hot part of a Bunsen flame.
- The colour of the flame identifies the positive ion (cation).

▼ **Table 14.3 Flame test results**

Positive ion (cation)	Flame colour
Lithium, Li^+	Red
Sodium, Na^+	Yellow
Potassium, K^+	Lilac
Calcium, Ca^{2+}	Orange–red
Barium, Ba^{2+}	Light green
Copper(II), Cu^{2+}	Blue–green

Using aqueous ammonia



▲ **Figure 14.8 Testing for cations (positive ions) in aqueous solution using aqueous ammonia**

Anions

Testing for halides (chloride, bromide and iodide), carbonates and nitrates was covered in Section 8.5. There are two other anions (negative ions) you should know how to identify. The tests and their results are shown in Table 14.4.

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▼ **Table 14.4 Testing for sulfites and sulfates**

Test	Result	Anion
Add acidified aqueous potassium manganate(VII)	Colour change from purple to colourless	Sulfite, SO_3^{2-}
Add dilute nitric acid, followed by aqueous barium nitrate	White precipitate	Sulfate, SO_4^{2-}

Gases

Tests for gases and their results are shown in Table 14.5.

▼ **Table 14.5 Testing for gases**

Test	Result	Gas
Damp red litmus paper	Turns blue	Ammonia, NH_3
Limewater	Turns milky	Carbon dioxide, CO_2
Glowing splint	Lights	Oxygen, O_2
Burning splint	Pops	Hydrogen, H_2
Damp litmus paper	Bleached	Chlorine, Cl_2
Acidified aqueous potassium manganate(VII)	Changes from purple to colourless	Sulfur dioxide, SO_2

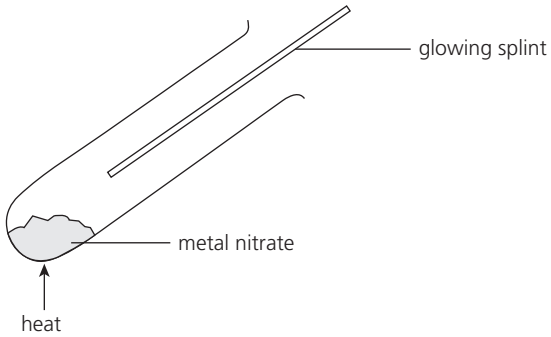
Revision activity

Use the information in this chapter and in Chapters 8 and 10 to create a branching flow chart to show how you would identify the ions present in an unknown solid. You could write the names of the tests on sticky notes and arrange them on a wall to help you decide the best order before you start to draw your chart.

Sample questions

REVISED

1 When some metal nitrates are heated, oxygen gas is given off. Oxygen relights a glowing splint.



A student heats four nitrates separately using a Bunsen burner. The student measures the time taken for the glowing splint to relight.

a It is important to ensure that the amount of heat produced by the Bunsen burner is kept constant in each experiment.
Suggest two methods of making sure the amount of heat is kept constant. [2]

b State another variable, concerning the metal nitrate, that should be kept constant to ensure that this is a fair test. [1]

c Name a piece of apparatus used to measure the time taken for the glowing splint to relight. [1]

d **i** Each experiment is repeated two more times and the results are shown in the table below.

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Metal nitrate	Time taken for glowing splint to relight/s			
	Experiment 1	Experiment 2	Experiment 3	Average time
Lithium nitrate	100	150	100	
Potassium nitrate	150	300	320	
Rubidium nitrate	400	450	410	
Sodium nitrate	200	200	240	

Use the times in the table to calculate an average result for each metal nitrate. Do not use any anomalous times. [4]

ii Name the metal nitrate that decomposes the fastest. Use the information from the table to explain how you made your decision. [2]

Student's answers

a 1 Use the same Bunsen burner in each experiment.
2 Put a thermometer into the solid.

b The concentration of metal nitrate should be the same.

c stopwatch

d i

Metal nitrate	Time taken for glowing splint to relight/s			
	Experiment 1	Experiment 2	Experiment 3	Average time
Lithium nitrate	100	150	100	116.67
Potassium nitrate	150	300	320	256.67
Rubidium nitrate	400	450	410	420
Sodium nitrate	200	200	240	213.33

ii Rubidium nitrate decomposes the fastest. The average time is the largest.

Teacher's comments

a Using the same Bunsen burner is a good start to an answer but more needs to be stated, as shown in the correct answers below. Monitoring temperature by putting a thermometer in the solid – or even the flame – is a common wrong answer to this type of question. The maximum reading on many laboratory thermometers is 110°C, so they should not be used for anything other than aqueous solutions unless you are told it is safe. High temperatures will cause the bulb to break, which is extremely hazardous if the liquid in the thermometer bulb is (toxic) mercury or a (flammable) organic liquid, such as ethanol.

b Metal nitrates are solids. Thus, concentration is not an option (it is only used for solutions). The volume of solids is not something that is normally measured.

c Any suitable timing device is acceptable.

d i The student has used *all* the results rather than just those that are either the same or close together. This illustrates the most important piece of advice for any exam: READ THE QUESTION CAREFULLY.

ii The longest time does not represent the fastest rate (see Chapter 7).

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Correct answers

a 1 Open the air hole of the Bunsen burner by the same amount in each case.
2 Make sure that the distance between the flame and the test-tube is the same in each case.

b mass of nitrate or number of moles of nitrate

c stopwatch

d i

Metal nitrate	Time taken for glowing splint to relight/s			
	Experiment 1	Experiment 2	Experiment 3	Average time
Lithium nitrate	100	150	100	100
Potassium nitrate	150	300	320	310
Rubidium nitrate	400	450	410	405
Sodium nitrate	200	200	240	200

ii Lithium nitrate. The shortest time represents the fastest rate of reaction.

2 A blue solid, **X**, contains one cation and one anion. The solid is dissolved in water and the following tests are carried out.

	Test	Observation
1	Add aqueous sodium hydroxide until it is in excess	Light blue precipitate, insoluble in excess
2	Add aqueous ammonia until it is in excess	
3	Acidify with dilute nitric acid, then add aqueous silver nitrate	No change
4	Acidify with dilute nitric acid, then add aqueous barium nitrate	White precipitate

a The solid is coloured. State what this suggests about the compound.

b State the conclusion that can be drawn from:
i Test 1
ii Test 3
iii Test 4

c State the observations you would expect to make in Test 2.

Student's answers

a *A transition metal is present.*

b i *Copper ions are present.*
ii *Cl^- ions are absent.*
iii *Sulfate ions are present.*

c *It forms a blue solution.*

Teacher's comments

a Solid **X** is a compound. Therefore, there could not be a transition metal present.

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b **i** You are expected to give the charge/oxidation state of the ion.
ii Acidification with dilute nitric acid, followed by aqueous silver nitrate, is a test for Cl^-, Br^- and I^-.
iii The student's answer is correct.
c This answer does not give detail about observations both before and after the ammonia is in excess.

Correct answers

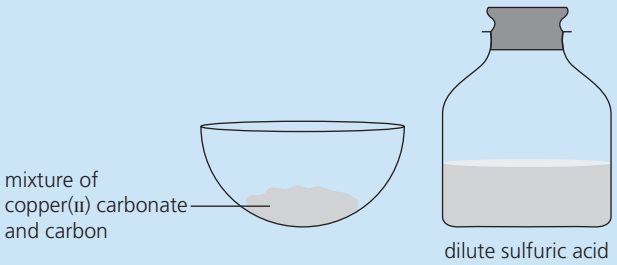
a Solid **X** contains ions of a transition metal.
b – Solid X contains Cu^{2+}.
– Solid X does not contain Cl^-, Br^- or I^-.
– Solid X contains SO_4^{2-}.
c Light blue precipitate, soluble in excess, giving a dark blue solution.

Exam-style questions

1 State the name of the process or processes that you would use to obtain:
a sugar crystals from a mixture of sugar and sand [3]
b pure water from an aqueous solution of copper(II) sulfate [1]
c liquid octane (boiling point 126°C) from a mixture of liquid octane and liquid decane (boiling point 174°C) [1]
d pure silver chloride from the precipitate formed when aqueous silver nitrate is added to dilute hydrochloric acid [3]
In some cases, only one process is required, but others may require more than one process. Explain your answer in each case. [Total: 8]

2 A student was told to make pure crystals of copper(II) sulfate from an aqueous solution of copper(II) sulfate. Describe how the student should carry this out. [Total: 4]

3 You are provided with a mixture of carbon and copper(II) carbonate.
Both substances are solids. Both solids are insoluble in water. Copper(II) carbonate reacts with dilute sulfuric acid and forms an aqueous solution. Carbon does not react with or dissolve in dilute sulfuric acid.



Use this information to plan an experiment to produce a sample of *pure carbon* from the mixture. [Total: 5]

4 A solid mixture, **R**, contains two cations and one anion.
The table below shows the tests a student does on an aqueous solution containing **R**.

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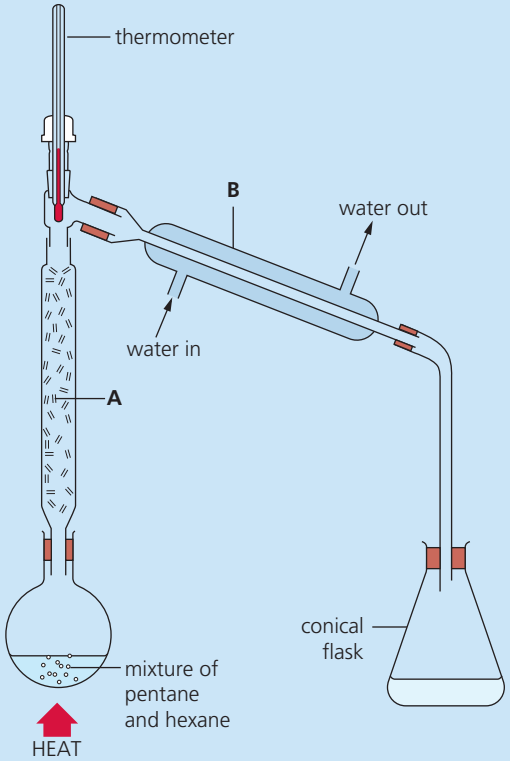
Test	Observation	Conclusion
Aqueous ammonia is added	[1]	**R** contains Cr^{3+} or Fe^{2+}
An excess of aqueous ammonia is added	[1]	**R** contains Cr^{3+} or Fe^{2+}
Aqueous sodium hydroxide is added	[1]	**R** contains Cr^{3+} or Fe^{2+}
Excess aqueous sodium hydroxide is added	[1]	**R** contains Cr^{3+}
The mixture from row above is warmed and the gas given off is tested with damp red litmus paper	[2]	[1]
[2]	[1]	**R** contains sulfate ion

Complete the table.
Identify any gases that are formed in the tests. [Total: 10]

5 A mixture of pentane and hexane is separated using the apparatus shown below.



a Name the pieces of apparatus labelled **A** and **B**. [2]
b Identify two errors in the apparatus. [2]
The errors are corrected and the separation is started.
c Name the process used to separate the mixture of liquids. [1]
d State why a Bunsen burner should not be used to heat the mixture of pentane and heptane. [1]
e Suggest why pentane collects in the conical flask before hexane. [1]
[Total: 7]

6 A student is given a mixture of two amino acids. The amino acids are both colourless solids that are soluble in water. Give full experimental details of how you would separate and identify the amino acids present in the mixture using paper chromatography. You are provided with all the necessary apparatus and a suitable locating agent. [Total: 6]

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Note: page numbers in **bold** refer to the location where a key definition is *first* defined.

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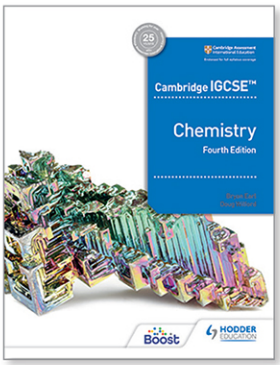

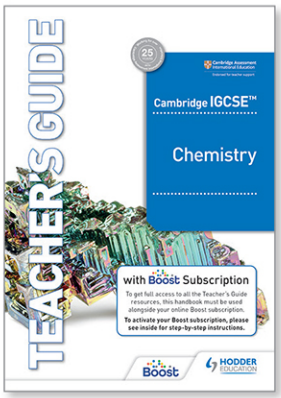

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